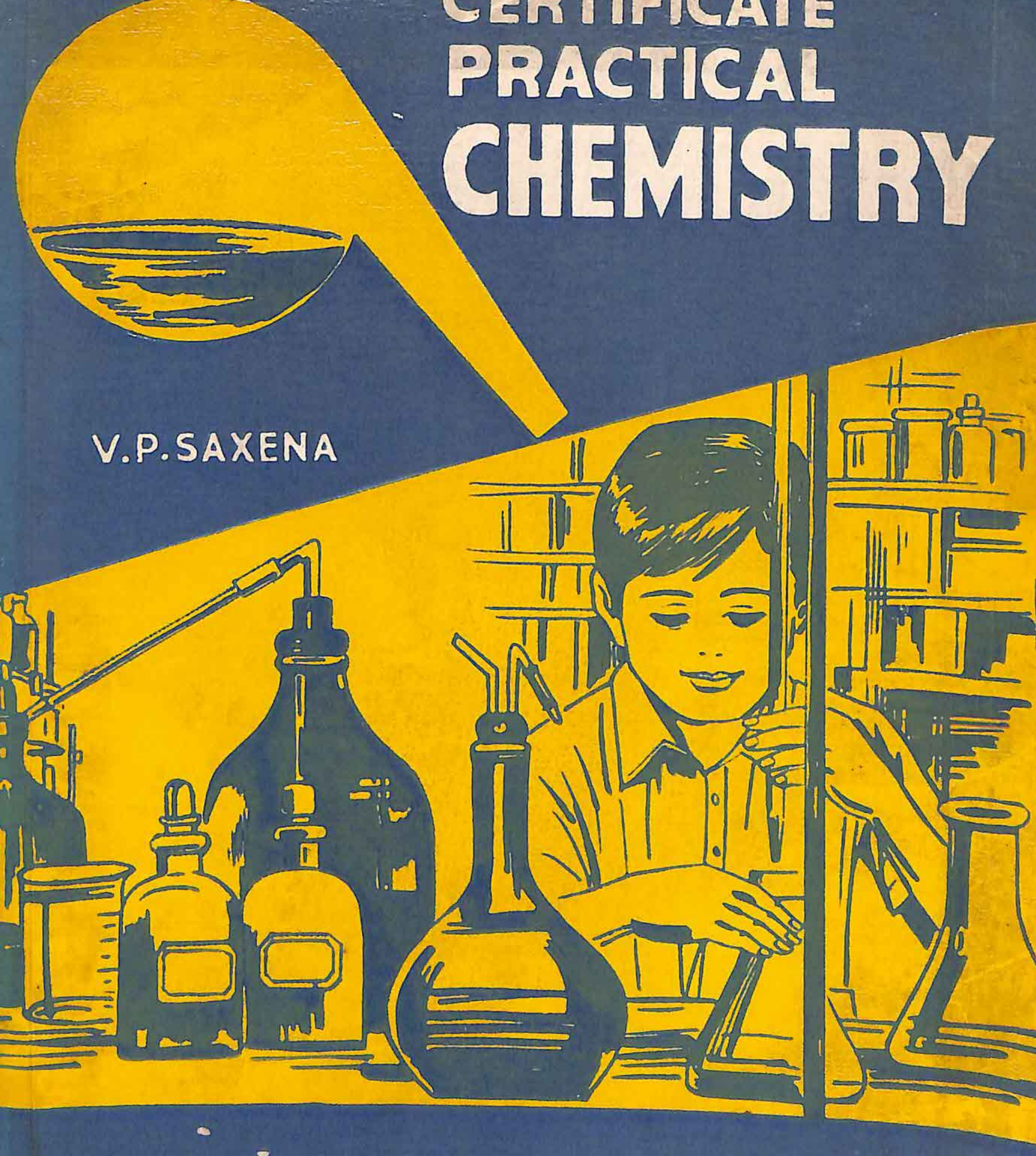
CERTIFICATE PRACTICAL CHEMISTRY

V. P. SAXENA



PITAMBAR BOOK DEPOT

Strictly based on the new syllabus in Practical Chemistry for I.C.S.E. Examination prescribed by the Council for I.S.C. Examinations, New Delhi.

CERTIFICATE PRACTICAL CHEMISTRY

(For Classes IX & X)

PARESH SAXENA

&

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PREFACE TO THE SECOND EDITION

The widespread acceptance of the first edition of *Certificate Practical Chemistry* has been highly gratifying. The thoroughly illustrated and explanatory descriptions of the experiments presented in it evoked much appreciation from teachers and students alike.

The present edition has provided an opportunity to rewrite all topics and reorganize them in a more logical sequence. More details have been provided wherever necessary. The most outstanding feature of the new edition is the large number of worked out exercises such as are given at the examination. These exercises will not only help student develop the ability to carry out the experiments methodically and to make correct observations and logical deductions, therefrom, but also to describe the experimental work done in a proper manner. We hope the book, in its present form, will be still more useful and enjoy wider popularity than before.

We gratefully acknowledge the many helpful comments of teachers who taught their students with the first edition, and look forward to receiving more comments and suggestions for further improvement of the book.

PARESH SAXENA
VED PRAKASH SAXENA

PREFACE TO THE FIRST EDITION

"*Certificate Practical Chemistry*" is designed to cover the requirements of the syllabus in practical chemistry prescribed by the *Council for the Indian School Certificate Examinations, New Delhi* for the ICSE examination. Since this book is intended to provide the first introduction to experimental work in chemistry, special efforts have been made to develop familiarity with apparatus, to train in careful experimentation, to generate a real feeling for the practical side of chemistry, and to illustrate the facts learnt in the theoretical studies of the subject. Detailed instructions along with diagrammatic illustrations have been provided for doing experiment so that the student should be able to collect and assemble the necessary apparatus and carry out the work competently.

The theoretical aspects involved in various experiments have been fully explained in a simple, straightforward language, since laboratory work done without a clear understanding of the underlying principles is only mechanical in nature and does not serve much useful purpose.

In spite of my best efforts, some errors might have escaped notice. I shall feel highly obliged if the teachers and students of the subject point out any errors and also send me their valuable suggestions to guide the development of this book through subsequent editions.

VED PRAKASH SAXENA

SYLLABUS IN PRACTICAL CHEMISTRY FOR I.C.S.E. EXAMINATION

1. Heat the given (unknown) substances, make observations, identify any products and make deductions where possible.
   - (*a*) mercuric oxide, red lead, lead dioxide
   - (*b*) copper corbonate, zinc carbonate
   - (*c*) washing soda, copper sulphate crystals
   - (*d*) zinc nitrate, copper nitrate, lead nitrate
   - (*e*) ammonium chloride, iodine, ammonium dichromate
2. Add concentrated sulphuric acid to the unknown substance, warm the mixture, make observations, identify the product by two tests and make deductions.
   (*c*) a chloride (*b*) a nitrate
3. Add dilute sulphuric acid to the unknown substance, warm if necessary, make observations, identify the product and make deductions.
   (*a*) a sulphide (*b*) a carbonate (*c*) a metal
4. Apply the flame test to identify the metal in the unknown substance, (*a*) a sodium salt, (*b*) a potassium salt, (*e*) a calcium compound.
5. Make a solution of the unknown substance; and sodium hydroxide-solution, make observations and give your deduction. Warming the mixture may be needed.

   Choose from substances containing
   Ca^{2-} , Cu^{2+} , Fe^{2+} , Fe^{3+} , Pb^{2+} , Zn^{2+} $(NH_4)^+$.
6. Supply a solution of a dilute acid and alkali.
   Determine which is acidic and which basic, giving two tests for each.
7. Add concentrated hydrochloric acid to each of the given substances, warm, make observations, identify any product and make deductions.
   (*a*) copper oxide (*b*) manganese dioxide.
8. The percentage composition of a mixture of powdered salt and water-washed sand.

   (The experiment would test techniques in dissolving, filtering or decanting washing and weighing. It may be counted out as taking so much time. The weakness could be met be supplying a given weight of the mixture; also by choosing sand of such grain size that filtering or decanting will not be slow and yet not so large that separation of salt and sand cannot be done simply by sorting out mechanically the sand from the salt. The experiment took about 20 minutes using 10 gm mixture (4 gm sand, 6 gm salt).

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1

General Introduction

Chemistry is essentially an experimental science. Nearly all chemical knowledge has been acquired through centuries of painstaking and skilled practical work. The modern chemist still depends on experimental evidence to further his knowledge. As such, the laboratory work continues to be an essential and important part of any course of study in chemistry. It provides abundant opportunities to make a first hand study of a wide variety of chemical reactions and processes. It has a sporting interest because of the inherent element of exploration.

In order to fully enjoy the practical work in chemistry and to get the maximum benefit out of it, you must first know your laboratory—the equipment in it, how to use it properly, and master certain basic skills essential to becoming a good experimenter.

1 The Chemical Laboratory

The chemical laboratory is the place where experiments in chemistry are done. It is provided with suitable equipment needed for chemical experiments. On your first visit to the laboratory, the teacher will take you round to acquaint you with the general arrangement of the laboratory and tell you where to find what. He will also instruct you how to handle the common equipment. You should make a careful note of all that is provided at your seat on the *working table* and elsewhere in the laboratory.

The working table is provided with—

1. *a gas tap* for the supply of gas to the burners,
2. *a water tap* for the ready supply of water,
3. *a sink*, for the disposal of liquid wastes,
4. *a waste box* (on the floor) for discarding solid wastes, and
5. *a reagent shelf* provided with a set of reagent bottles containing reagents of more common use,

Elsewhere, you will find—

1. *a side shelf* fixed on the wall, on which are placed chemicals and reagents of less frequent use,
2. *a gas plant* for producing the gas (generally, petrol gas) for burning in the burners,
3. *a water still* for producing distilled water required in many experiments, and
4. *the demonstration table* on which the teacher performs experiments for demonstration to a whole group of students

Near the roof of the laboratory you will find the exhaust fans which work to exhaust polluted air from the laboratory.

2. The Proper Use of Laboratory Equipment

In the use of all laboratory equipment, you must always bear two points in mind: economy and safety. Do not be wasteful with gas, water, or chemicals.

1. Gas tap. Use the gas with utmost economy. If the gas supply is good the burner flame may be bigger than is necessary when the tap is turned full on. In that case turn the tap to reduce the flame to a proper size. Turn off the gas when it is no longer required. Never allow the gas to leak into air; it is both wateful and harmful.

2. Water tap. Notice the arrangement of water taps and the sink. When required, turn the tap to get water in a gentle stream, and then turn off. A vigorous stream of water may cause splashing on your working table. *Never use the water from the laboratory taps for drinking.*

3. Sink. Use the sink for throwing waste liquids only. Do not throw any solid materials, *e.g.*, and filter papers, broken pieces of apparatus, etc. into a sink; this may result in the blocking of the outlet pipe causing much inconvenience. Neve. throw hot concentrated acids directly into the sink. Allow to cool, and then pour into running water.

4. Waste box. Waste box is the recepticle for solid wastes, *e.g.*, used filter papers, broken glass or porcelain, used splints, etc. Do not let such things lie about on your working table and do not throw them about on the floor. Before throwing away a lighted splint or match stick, extinguish it and make sure that it has no spark left.

5. Reagent shelf. The reagent shelf on your working table is provided with a number of bottles containing reagents which are used very frequently. These reagent bottles are arranged in a definite order. The following instructions should be kept in mind with regard to the use of the reagent bottles:

(*i*) Take out only one reagent bottle at a time, and return it to its proper place immediately after use. If too many bottles are taken out at a time, their locations are likely to be upset through haphazard replacement. Moreover, their stoppers might also get mixed up causing the reagents to be contaminated.

(*ii*) Never put down a reagent bottle or its stopper on the bench. Directly pour the reagent from the bottle into a clean test tube, beaker, or any other vessel, holding its stopper in a suitable manner.

(*iii*) Never dip a spatula, glass rod, filter paper strip, or any other transferal device to remove reagents from their bottles, since this practice may contaminate the reagents.

(*iv*) Do not measure out a large excess of a reagent from its bottle since chemicals are usually expensive. However, any excess reagent measured out by negligence must not be returned to its bottle since this is another possible source of contamination. It should be discarded.

(*v*) Transfer solid reagents with a spatula into the clean receiving vessel. Never transfer it on your palm which is unsafe or on a filter paper which would be wasteful. *Never pick up chemicals in your fingers unless you are told to do so.*

6. *Side shelf*. Never remove a bottle from the side shelf to your seat. The reagents on the side shelf are meant for the collective use of the entire group of students working at a time in the laboratory.

3. Tips for Good Work

1. Read the experiment to be done in advance and thoroughly understand the procedure.

2. Never forget to bring with you to the laboratory—
   - (*i*) the text book for on the spot consultation,
   - (*ii*) the practical notebook for entering results, and
   - (*iii*) two clothes for cleaning purposes : a rough cloth to wipe up any liquid spilt on the bench, and a limp cloth for drying the insides of test-tubes, beakers, etc.

3. Note all the intstructions given by your teacher with regard to the experimental work.

4. Check up before starting the experiment that everything you would need is readily available. If you find a reagent bottle empty, do not remove it from the seat of another student; ask the laboratory staff to refill it. Also make sure that the reagent solutions are clear and free from turbidity

5. Always use scrupulously clean apparatus; this would ensure good results. *It is impossible to do good work with dirty apparatus.*

6. Before starting the experiment, arrange all the things required at the rear of the table keeping the space in front of you clear and uncluttered.

7. Do your work honestly, confidently, and patiently. In case of any difficulty, whatsoever, approach your teacher for guidance. Do not indulge in loose talks with fellow students or the laboratory staff.

8. At the conclusion of the laboratory session—
   - (*i*) thoroughly clean the work space and glassware that has been used,
   - (*ii*) put away everything in your locker, and secure it with a good padlock, and
   - (*iii*) make sure that the gas and water taps have been turned off.

4. Safety Precautions

The possibility of accidents happening in chemical laboratories can never be ruled out. However, whenever an accident happens, it is nearly always because some elementary precautions have been neglected. The following safety precautions must be observed by a student while working in the chemical laboratory :

1. It is advisable to wear a lab coat or apron to protect clothing.

2. Read the labels on all chemical reagent containers before using their contents. The use of a wrong reagents can lead to a serious accident.

3. Always wipe up quickly liquids that are spilt on the bench. Acids or bases spilt on the bench or floor should be neutralized with solid sodium bicarbonate before removal.

4. If an acid spills on clothing, neutralize it with a very dilute ammonia or sodium bicarbonate solution. When bases are spilt, first neutralize with a dilute acetic acid solution and then wash the affected part with dilute ammonia solution.

5. Wash immediately any solids or liquids on skin or clothes with plenty of water. The absorption of a chemical through the skin could be a potential mode of poisoning. Never forget to wash the hands just prior to leaving the laboratory for the day.

6. If a chemical enters the eyes or mouth, rinse it out with a large amount of water at once, and report the accident to your teacher. He will arrange the required first aid.

7. Never taste a chemical or inhale its vapour unless your teacher instructs you to do so. Always assume that a particular chemical is hazardous unless you know for sure that it is not.

8. Never use cracked or broken glassware; it may cause an injury or at the very least, fall apart during the course of an experiment. Replace any damaged glassware at once.

9. Always allow a hot test tube to cool before placing it in the rack, if the rack is wet.

10. Always carry out an experiment exactly according to directions. Never add a reagent in excess of the quantity you have been told to use. *Never carry out unauthorized experiments.*

Additional safety precautions are given along with the various laboratory techniques in chapter 2.

2

Common Laboratory Techniques

There are many simple laboratory techniques and basic operations which find common application in all kind of experimental work in chemistry. Unless these techniques and basic skills are mastered thoroughly, you should not hope to become a good experimenter. This chapter describes some of the common laboratory techniques in detail, such as using the Bunsen burner, handling glass apparatus, handling chemicals and reagents, heating substances, filtration, evaporation, etc.

1. Using the Bunsen Burner

Take a Bunsen burner. Unscrew its long metal tube, called the chimney or *barrel* and examine carefully its various parts. You will find that it consists of three parts : the base-plate, the barrel, and the air regulator.

1. The base-plate. It is a heavy disc-like piece of cast iron. Its function is to provide support to the entire structure of the burner.

The base-plate has a small, horizontal side tube, called the *gas tube*. It is connected by means of rubber tubing to the gas tap on the bench.

At the top of the base-plate is a pin-point jet, called, *nozzle*. The nozzle communicates with the gas tube through a very fine hole running throughout its length. When the gas tap on the bench is turned on, the gas comes along the rubber tubing, enters the gas tube, and then issues at the nozzle under increased pressure.

2. The barrel: This is a long metal tube which screws over the nozzle on top of the base-plate. Note that there is a large hole near its bottom opposite the nozzle. This hole is called the *air-hole*. Its function is to admit air. As the gas rushes up the barrel from the nozzle, air is sucked in through the air-hole. The mixture of the gas and air passes upward and is lighted at the top of the barrel by applying a flame.

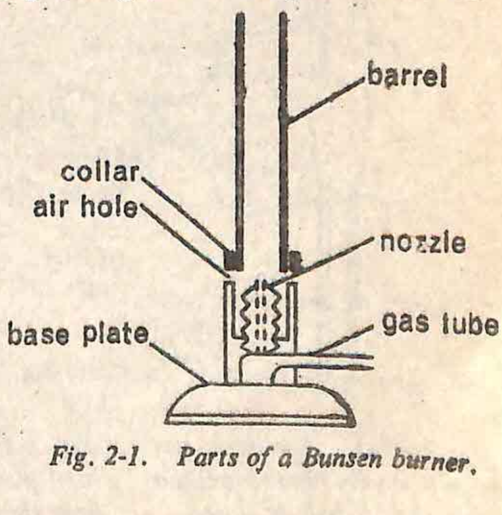

Fig. 2-1. Parts of a Bunsen burner.

3. *The collar.* This is a metal ring that can be turned round the lower end of the barrel. It has a hole of the same size as the hole in the barrel. By turning this ring or the collar, the air hole in the barrel can be closed or opened to any desired extent. Its function, therefore, is to regulate the amount of the air mixing with the gas.

How to light the Bunsen burner ? The Bunsen burner should always be lighted in the following way :

1. Close the air-hole completely.

2. Turn the gas on full.

3. Apply a light at the mouth of the barrel to ignite the gas. A long, yellow, smoky, and flickering flame is produced (Fig. 2-2 *a*). It is not suitable for heating because it is not very hot and it deposits soot on the vessel being heated.

4. Turn the gas down until the flame is reduced to about 8-10 cm length.

5. Move the collar slowly to adjust the air-hole untill the flame becomes non-luminous (Fig. 2-2 *b*).

If the air-hole is opened too much, the flame may 'strike back', *i.e.*, travel downward the burner tube and ignite the incoming gas at the nozzle. Thus, instead of a normal flame at the top of the barrel, a tiny flame is produced at the nozzle accompanied with a hoarse whistling noise (Fig. 2-2 *c*). Should the flame strike back, turn off the gas at the tap, allow the burner to cool, and relight after adjusting the air-hole and the flow of gas properly.

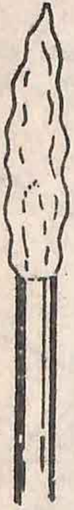

Fig. 2-2 (a). A yellow, unsteady, smoky flame produced when the air-hole is closed.

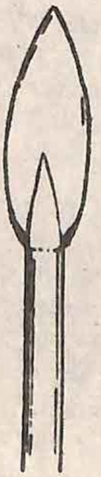

Fig. 2-2 (b). A steady, non-luminous flame with the air-hole rightly open.

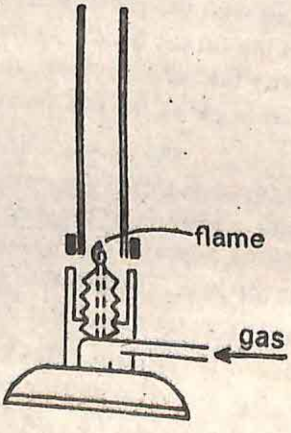

Fig. 2-2 (c). Striking back of a Bunsen flame when the air-full is too-much open.

2. Handling Glass Apparatus

1. Usually the glass apparatus used in chemical experiments are fragile, have thin walls, and are delicate. They should, therefore, be handled with care. Slap dash manner of handling may cause the glass apparatus to break during an experiment and thereby spoil it.

2. Never place a hot apparatus on a cold surface; always place it on an asbestos wire gauze or on an asbestos pad.

3. Thick walled glass apparatus, *e.g.*, reagent bottles, graduated cylinders, etc. must never be heated, nor should hot liquids be poured into them. They are liable to crack on heating.

4. Thin test tubes may be heated directly in a flame; breakers and flasks should be placed on an asbestos wire gauze resting on a tripod or ring. The abestos ensures even distribution of heat over the bottom of the vessel.

5. Never use a glass apparatus for any other purpose than what it is meant for.

3. Handling Chemicals and Reagents

Solids. A solid material is usually stored in a wide mouth bottle. The best way to transfer the solid is to tilt the bottle and rotate it to and fro unil the required quantity of the solid has been transferred directly to another container (*e.g.*, a watch glass, beaker, etc.). If at all you have to use a spatula, make sure that it is absolutely clean. If a solid is not easily pourable, loosen it by tapping the container and gentle shaking. Never attempt to break up a caked solid inside the bottle with an instrument.

Liquids. When pouring a liquid from a bottle, proceed as follows :

(*i*) Read the label before you pour to make sure that you are using the right reagent.

(*ii*) Remove the stopper from the bottle by holding the edge of the stopper between the first and second fingers of the right hand, having the palm up. Reversing the hand, grasp the bottle with its label facing upward. In this position the label would not be damaged by liquid which might run from the neck down the outside of the bottle when it is put down.

(*iii*) Pour the liquid gently along the sides of the receiving vessel held in the left hand at a proper slant (Fig. 2-3). Poured out in this manner, the liquid does not splash.

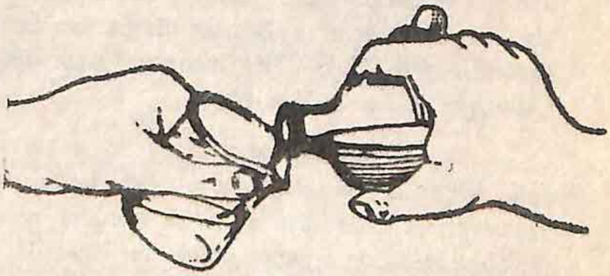

Fig. 2-3. The correct way to pour a liquid from a bottle.

(iv) After the required quantity of the liquid has been transferred, remove the drop of the liquid adhering to the tip by touching it to the tip of the receiving vessel. Replace the stopper, and put the bottle back on the shelf.

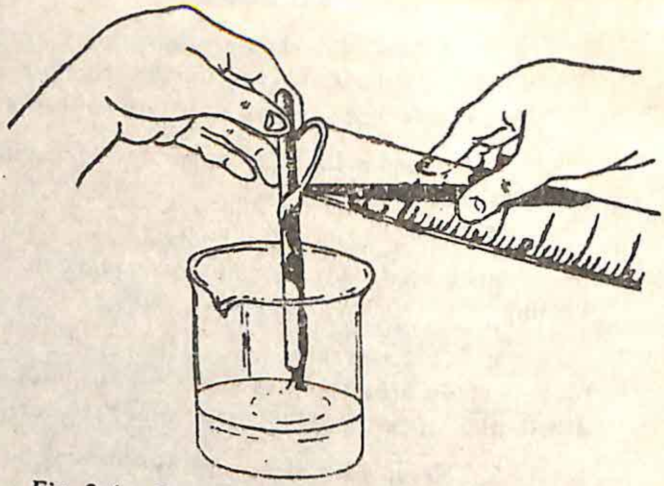

Fig. 2-4. Pouring liquid from a measuring cylinder into a beaker and directing the flow with a glass rod.

Although liquids can often be poured directly from one vessel (*e.g.*, a beaker, measuring cylinder, etc. without special precautions, it is good practice to direct the flow with a clean glass rod (Fig. 2-4).

Never poke chemicals, solids or liquids, with a pen, pencil, etc. This practice results in contamination of the chemicals, and is extremely unsafe. Use a piece of glass rod to poke a solid or stir a liquid.

4. Heating Substances

In a non-luminous Bunsen flame, the temperature prevails highest a little below the tip of the flame and drops sharply down the middle towards the base. Hence, in order to fully utilize the heat of the flame, the object (*e.g.*, a test tube) should be located a little below the tip of the flame, and not in middle part. Hence, the correct way of heating in a Bunsen flame is as shown in Fig. 2-5. The incorrect way of heating is shown in Fig. 2-6.

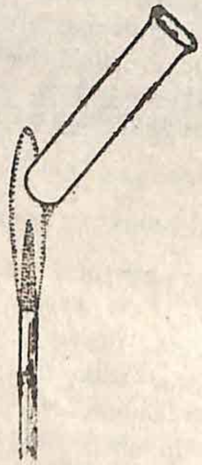

Fig. 2-5. Correct way of heating on a Bunsen flame

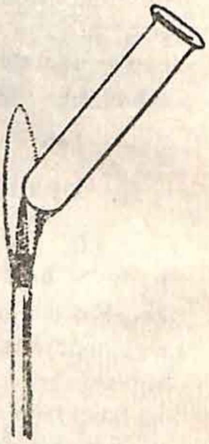

Fig. 2-6. Incorrect way of heating on a Bunsen flame.

When heating something in a test tube, hold the test tube by a test tube holder. However, you can have a better control of the tube by using a paper collar instead. It can be made by folding a sheet of paper lengthwise in half, and then in half again so as to have a thick paper strip about 25 cm long and 2·5 cm wide. Place it round the top end of the test tube, and firmly hold the two free ends between the thumb and the forefinger pressing against the tube. In this manner, the tube is gripped tightly inside the paper collar and can be held in the flame without risk of burned fingers.

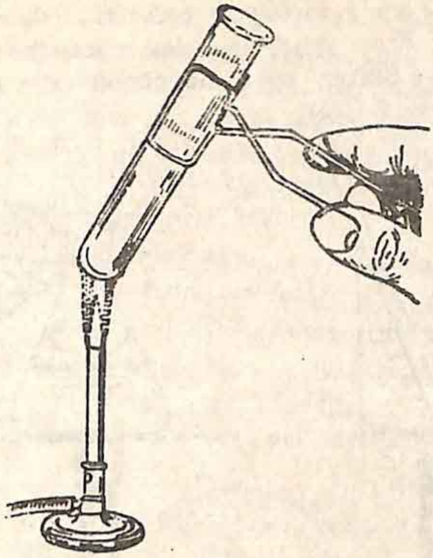

Fig. 2-7. *Holding the test tube by a test tube holder.*

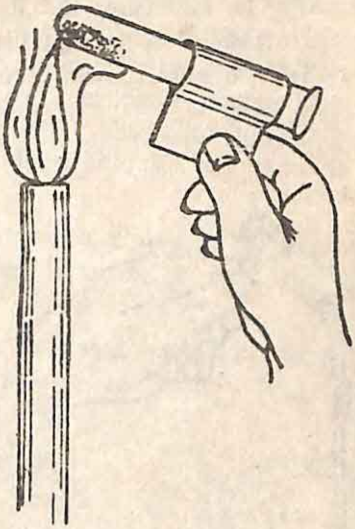

Fig. 2-8. *Holding the test tube by a paper collar.*

When heating the tube, ensure that its open end is pointing away from you and from anyone else who may be near you. If possible, the tube should be pointed towards the wall. Never look down the length of a test tube to observe the contents. These precautions will avoid mishap in case the contents happen to be violently ejected from the tube in the course of heating (Fig. 2-9).

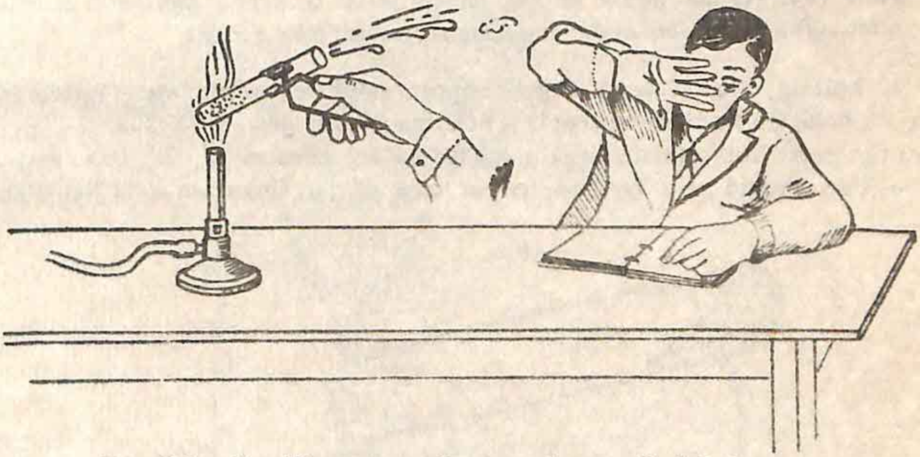

Fig. 2-9. *Handling a test tube towards oneself while heating may cause a mishap.*

During heating, a small region of the liquid may get superheated and suddenly vaporize, causing the liquid to splash or 'bump'. To prevent this, always move the tube back and forth slightly but continually in the flame and occasionally shake the contents. The flame should be allowed to strike a point just below the liquid surface in order to reduce bumping [Fig. 2-10(*a*)]. In any case, the flame should never touch the glass above the level of the liquid in the test tube. If the liquid is heated in a beaker, stir it continuously with a glass rod [Fig. 2-10(*b*)]; if in a flask, swirl it constantly [Fig. 2-10(*c*)].

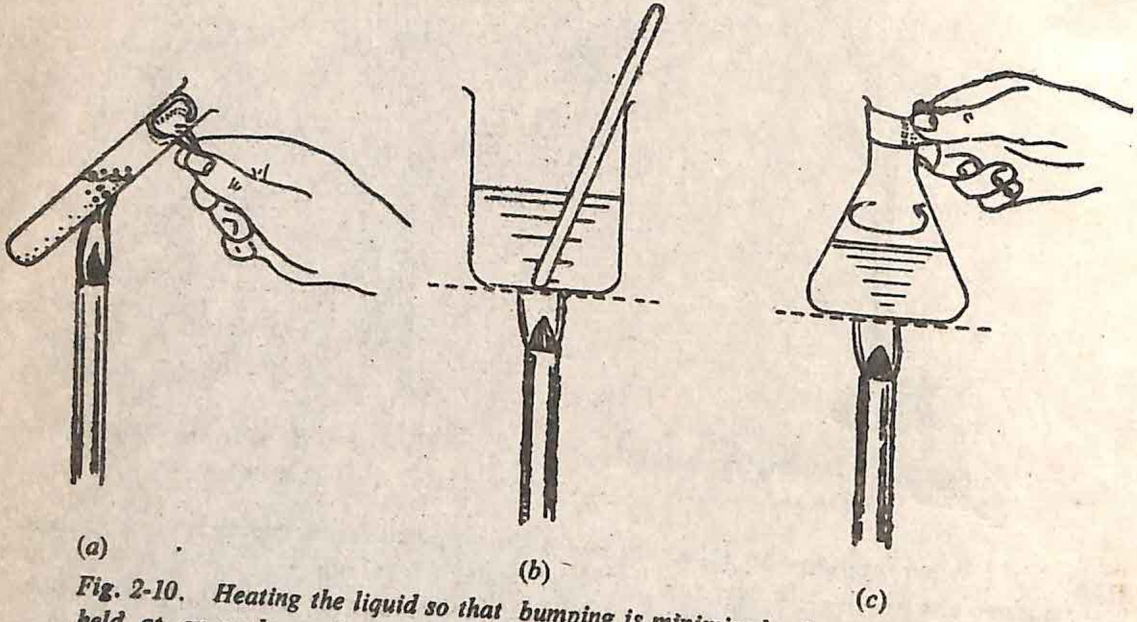

Fig. 2-10. Heating the liquid so that bumping is minimized: (a) the test tube is held at an angle, and moved back and forth through the flame directed towards a point just below the liquid surface, (b) the liquid in the beaker is stirred with a glass rod, (c) the liquid in the conical flask is kept in motion by swirling the flask. [The dotted line in the diagrams represents wire gauze.]

If heating a solid (*e.g.*, sugar, copper sulphate crystals, etc.), results in water being driven off, hold the test tube so that it is pointing slightly downward (Fig. 2-11). This would prevent the water from running back down the hotter portion of the tube and cracking it. This position should also be adopted in case of any unknown solid lest it should give off water.

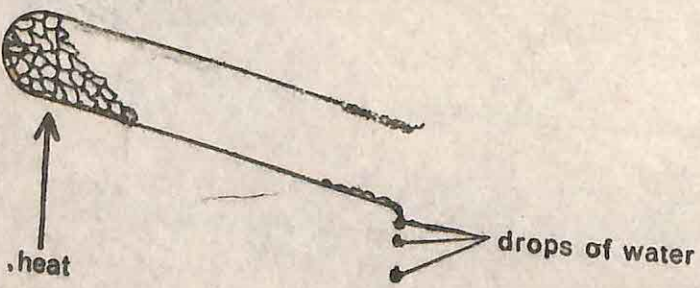

Fig. 2-11. Position of the test tube when heating a substance that gives off water.

Whatever substance you may be heating and whatever vessel you may be using, *heat should always be applied gently at first.*

5. Filtration

Filtration is the process employed for removing an insoluble substance from water or a solution of another substance. The insoluble substance obtained by filtration is called the *residue*; the liquid obtained is called the *filtrate*. In the laboratory, filtration is done through an ordinary funnel fitted with a *filter paper*, which is a specially prepared porous paper. Thus, the filter paper acts as a sieve, holding back the solid and allowing the liquid to pass through. The liquid passing through the filter paper drips through the stem of the funnel. For filtration, proceed as follows :

1. Fold a circular filter paper first into half and then into a quarter as shown in Fig. 2-12 (*a, b*). Now open the folded filter paper so that it forms a cone with three thicknesses on one side and single thickness on the other (Fig. 2-12 *c*).

2. Place the filter paper cone in the funnel and press it down with your thumb so that its upper edge fits snugly against the funnel walls. Moisten the cone with a little water when it will stick tightly to the glass, and is ready for use (Fig. 2-12*d*).

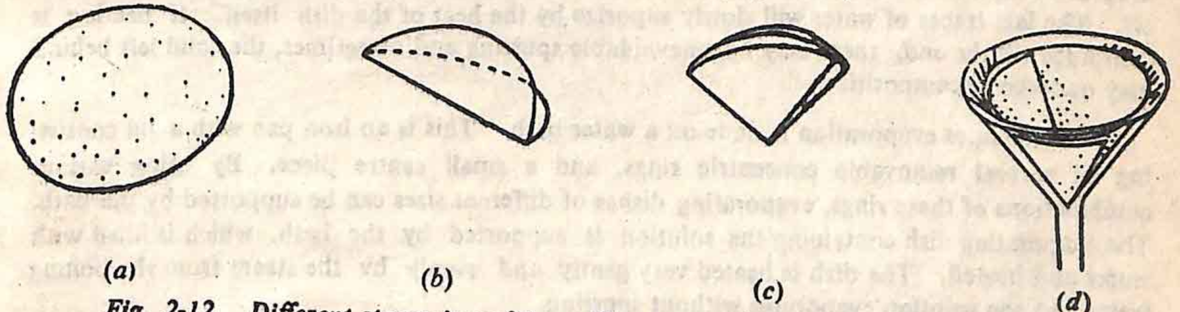

Fig. 2-12. Different stages in making a filter paper cone and mounting it in a funnel

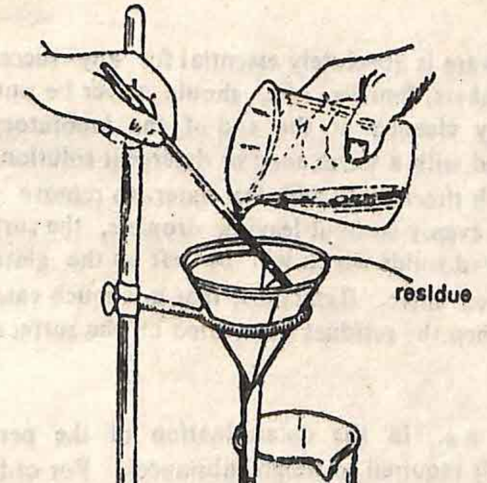

Fig. 2-13. Filtration through the filter paper cone.

3. Mount the funnel on a funnel stand. Place a beaker (or any suitable receiving vessel) under it, to collect the filtrate.

4. Pour the solid-liquid mixture to be filtered down a clean glass rod touching the filter paper, Fig. 2-13. If the quantity is large, start the filtration by decanting liquid into the funnel so that the pores of the filter paper are not clogged with the solid particles at the very start. Transfer the bulk of the solid along with the last portion of the liquid.

At no time the level of the liquid should rise more than two-thirds up the filter paper cone.

6. Evaporation

Evaporation is the process employed for the recovery of a solid from solution; the solvent being driven off as vapour. It is generally carried out by heating the solution in an evaporating dish on a gauze. For more gentle and even heating, a sand bath is used which is a shallow tray containing sand. The solution should be frequently stirred with a glass rod, to prevent spurting. If the solution is to be evaporated to 'dryness', do not heat till the last drop has evaporated but remove the source of heat after the bulk of the water has been driven off. The last traces of water will slowly vaporize by the heat of the dish itself. If heating is continued till the end, there may be unavoidable spurting and sometimes, the solid left behind may undergo decomposition.

Sometimes evaporation is done on a water bath. This is an iron pan with a lid consisting of several removable concentric rings, and a small centre piece. By using various combinations of these rings, evaporating dishes of different sizes can be supported by the bath. The evaporating dish containing the solution is supported by the bath, which is filled with water and heated. The dish is heated very gently and evenly by the steam from the boiling water and the solution evaporates without spurting.

7. Cleaning Glassware

The use of clean glassware is absolutely essential for any successful experimental work in chemistry. Test tubes, beakers, funnels, etc., should never be put away in a dirty condition, but should be thoroughly cleaned at the end of the laboratory work. The common glassware can be easily cleaned with a warm soap or detergent solution, using a test tube brush if necessary. Afterward, wash thoroughly with tap water to remove the detergent from glass surface. If the water drains evenly without leaving droplets, the surface is clean. Since tap water always contains dissolved solids which will be left on the glass when the water evaporates, rinse finally with distilled water. Remember that it is much easier to clean the glassware immediately after use than when the residues have dried on the surface.

8. Weighing

In many experiments, *e.g.*, in the determination of the percentage composition of mixtures (see Chapter 8), it is required to weigh substances. For ordinary weighing, one may use a trip balance (Fig. 2-14).

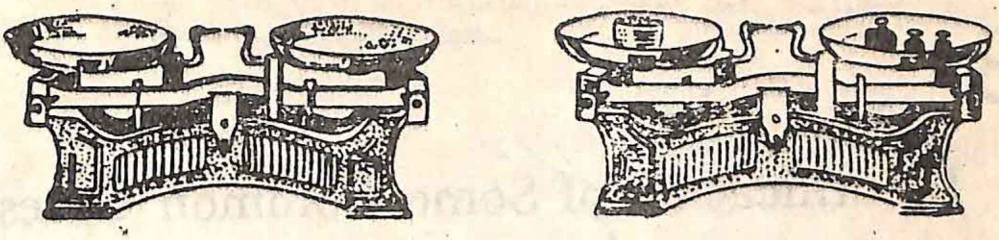

Fig. 2.14 (a). A trip balance *(b) Weighing in a trip balance*
(The article to be weighed is placed in the left pan and the weights on the right pan)

For more accurate weighing, however, a chemical balance is used. The chemical balance is a delicate and sensitive instrument, and an object can be weighed on it within decimal fractions of a milligram.

3

Identification of Some Common Gases

Identification of gases is not only an exercise valuable in itself, it is also an important step in the identification of many a substances. Most substances give off some gas or other when heated alone or when treated with certain reagents, *e.g.*, acids, alkalis, etc. The identification of a gas produced from a given substance by a particular reaction leads one step towards the identification of that substance. Thus, if it is found that a substance gives off ammonia when heated with an alkali, the substance surely is an ammonium compound. And, if the same substance is also found to give off hydrogen chloride when heated with concentrated sulphuric acid, its identity is completely established—it is ammonium chloride !

In this chapter we shall study the more common reactions in which some common gases (*viz.*, hydrogen, oxygen, carbon, dioxide nitrogen dioxide, chlorine, hydrogen chloride, sulphur dioxide, ammonia, and water vapour) are formed, and the tests by which each of them is identified.

1. Hydrogen, H_2

Hydrogen is formed when an active metal, *e.g.*, zinc, iron, etc., reacts with dilute sulphuric acid or dilute hydrochloric acid.

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$$
$$Fe + 2HCl \rightarrow FeCl_2 + H_2\uparrow$$

It is also formed when zinc or aluminium (in the form of shavings) is heated with a concentrated solution of an alkali—sodium hydroxide or potassium hydroxide.

$$Zn + 2NaOH \rightarrow \underset{\text{sodium zincate}}{Na_2ZnO_2} + H_2\uparrow$$
$$2Al + 2NaOH + 2H_2O \rightarrow \underset{\text{sodium aluminate}}{2NaAlO_2} + 3H_2\uparrow$$

Tests. (1) Hydrogen is a colourless and odourless gas.

(2) When a lighted splint is brought near hydrogen, it burns with hardly visible flame, making a squeaky pop.

Experiment 3-1. To test for hydrogen.

Place a small piece of zinc in a test tube and add a few cm³ of dilute hydrochloric acid. Effervescence takes place due to the formation of tiny bubbles of a gas.

1. Observe the colour of the gas; it is colourless.

2. Carefully smell the gas; it is odourless.

3. Bring a lighted taper near the mouth of the test-tube; there is a squeaky pop and the gas burns with hardly visible pale-blue flame.

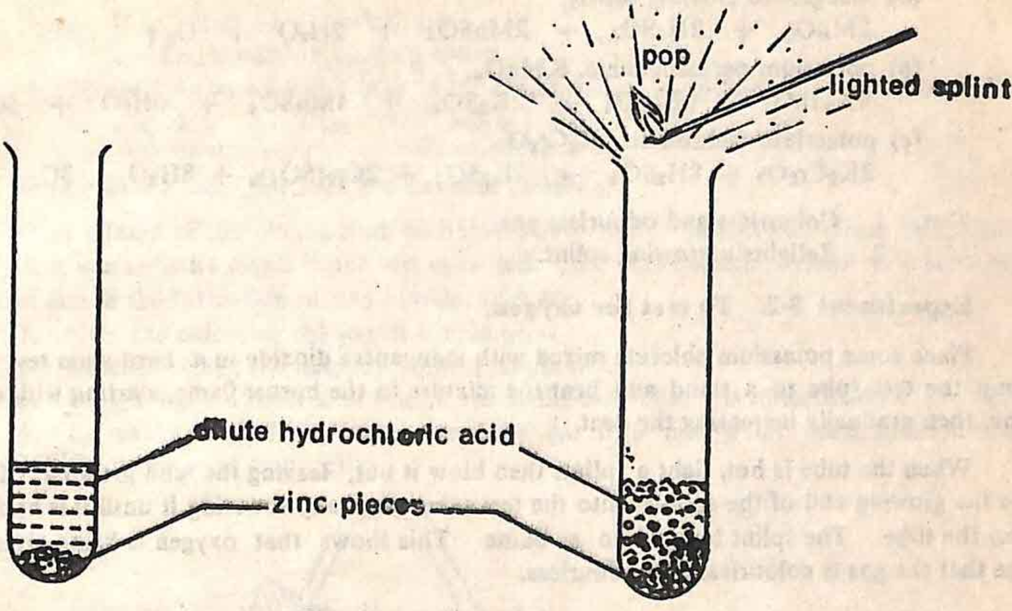

Fig. 13-1. Testing for hydrogen.

All these tests show that the gas is hydrogen.

2. Oxygen, O_2

Oxygen is formed in several reactions. For example—

1. When oxides of certain metals are heated, *e.g.*, mercuric oxide, red lead, etc.

$$\underset{\text{mercuric oxide}}{2HgO} \xrightarrow{\Delta} \underset{\text{mercury}}{2Hg} + \underset{\text{oxygen}}{O_2\uparrow}$$

$$\underset{\text{read lead}}{2Pb_3O_4} \xrightarrow{\Delta} \underset{\text{lead monoxide}}{6PbO} + \underset{\text{oxygen}}{O_2\uparrow}$$

2. When a chlorate, nitrate, permanganate, or dichromate of sodium or potassium is heated.

$$\underset{\text{pot. chlorate}}{2KClO_3} \xrightarrow{\Delta} \underset{\text{pot. chloride}}{2KCl} + \underset{\text{oxygen}}{3O_2\uparrow}$$

$$\underset{\substack{\text{sodium}\\\text{nitrate}}}{2NaNO_3} \xrightarrow{\Delta} \underset{\substack{\text{sodium}\\\text{nitrite}}}{2NaNO_2} + \underset{\text{oxygen}}{O_2\uparrow}$$

$$\underset{\substack{\text{potassium}\\\text{permanganate}}}{2KMnO_4} \xrightarrow{\Delta} \underset{\substack{\text{potassium}\\\text{manganate}}}{K_2MnO_4} + \underset{\substack{\text{manganese}\\\text{dioxide}}}{MnO_2} + \underset{\text{oxygen}}{O_2\uparrow}$$

$$\underset{\substack{\text{potassium}\\\text{dichromate}}}{4K_2Cr_2O_7} \xrightarrow{\Delta} \underset{\substack{\text{potassium}\\\text{chromate}}}{4K_2CrO_4} + \underset{\substack{\text{chromium}\\\text{trioxide}}}{2Cr_2O_3} + \underset{\text{oxygen}}{3O_2\uparrow}$$

3. When sodium peroxide is acted upon by water.

$$2Na_2O_2 + 2H_2O \rightarrow 4NaOH + O_2\uparrow$$

4. When concentrated sulphuric acid is heated with—
   (*a*) manganese dioxide, MnO_2,
   $$2MnO_2 + 2H_2SO_4 \rightarrow 2MnSO_4 + 2H_2O + O_2\uparrow$$
   (*b*) potassium permanganate, $KMnO_4$,
   $$4KMnO_4 + 6H_2SO_4 \rightarrow 2K_2SO_4 + 4MnSO_4 + 6H_2O + 5O_2\uparrow$$
   (*c*) potassium dichromate, $K_2Cr_2O_7$.
   $$2K_2Cr_2O_7 + 8H_2SO_4 \rightarrow 2K_2SO_4 + 2Cr_2(SO_4)_3 + 8H_2O + 3O_2\uparrow$$

Test. 1. Colourless and odourless gas.
2. Relights a growing splint.

Experiment 3-2. To test for oxygen.

Place some potassium chlorate mixed with manganese dioxide in a hard-glass test tube. Clamp the test tube to a stand and heat the mixture in the burner flame, starting with a low flame, then gradually increasing the heat.

When the tube is hot, light a splint, then blow it out, leaving the end growing. Introduce the glowing end of the splinter into the test tube, gradually lowering it until it is half way down the tube. The splint bursts into a flame. This shows that oxygen is being given off. Note that the gas is colourless and odourless.

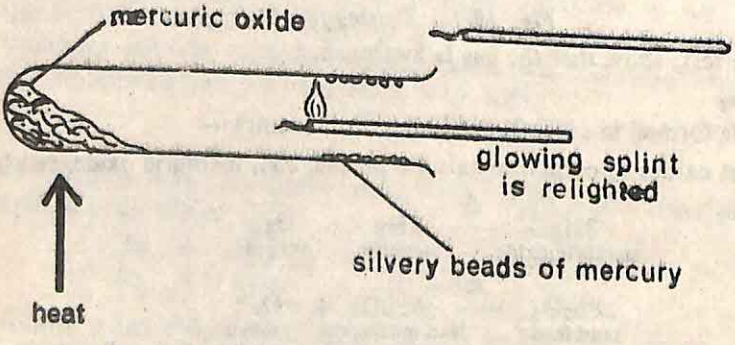

Fig. 3-2. The 'glowing splint test' for oxygen.

3. Carbon Dioxide, CO_2

Carbon dioxide is evolved when—

1. Metal carbonates (other than those of sodium and potassium) are strongly heated.

$$CuCO_3 \xrightarrow{\Delta} CuO + CO_2\uparrow$$

$$ZnCO_3 \xrightarrow{\Delta} ZnO + CO_2\uparrow$$

2. Sodium hydrogen carbonate is strongly heated.

$$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + H_2O + CO_2\uparrow$$

3. Metal carbonates or hydrogen carbonates are treated with dilute acids.

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2\uparrow$$
$$MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + H_2O + CO_2\uparrow$$
$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2\uparrow$$

Tests. 1. Colourless, odourless gas.

2. Extinguishes a burning splint.

3. Turns lime water milky (due to the formation of insoluble calcium carbonate).

$$\underset{\text{lime water}}{Ca(OH)_2} + CO_2 \rightarrow \underset{\text{calcium carbonate}}{CaCO_3\downarrow} + H_2O$$

Experiment 3-3. To test for carbon dioxide.

Place a lump of marble (calcium carbonate) in a test tube and add dilute hydrochlori acid (about a centimetre depth in the test tube will do). Immediately there is a brisk effe vescence due to the formation of tiny bubbles of a gas.

1. Note the colour of the gas; it is colourless.
2. Carefully smell the gas; it is almost odourless.
3. Bring a lighted taper near the mouth of the test tube; it is immediately extinguishe
4. By means of a delivery tube, pass the gas into lime-water taken in another te tube; the lime water soon turns milky.

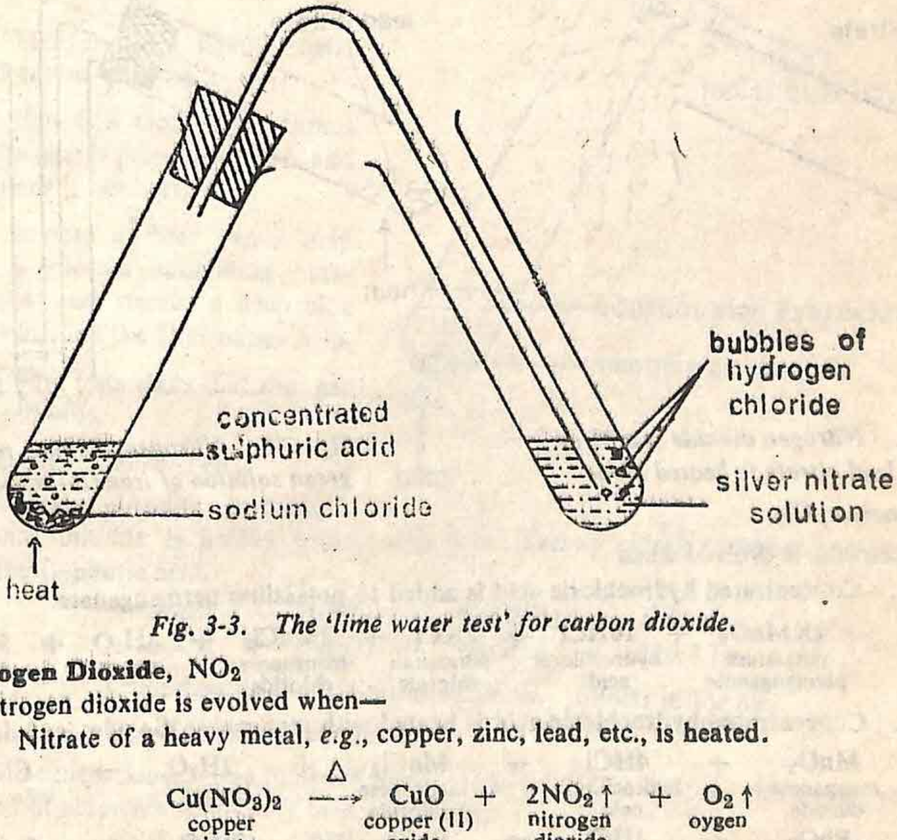

Fig. 3-3. The 'lime water test' for carbon dioxide.

4. Nitrogen Dioxide, NO_2

Nitrogen dioxide is evolved when—

1. Nitrate of a heavy metal, *e.g.*, copper, zinc, lead, etc., is heated.

$$\underset{\substack{\text{copper}\\\text{nitrate}}}{Cu(NO_3)_2} \xrightarrow{\triangle} \underset{\substack{\text{copper (II)}\\\text{oxide}}}{CuO} + \underset{\substack{\text{nitrogen}\\\text{dioxide}}}{2NO_2\uparrow} + \underset{\text{oygen}}{O_2\uparrow}$$

$$\underset{\text{lead nitrate}}{2Pb(NO_3)_2} \xrightarrow{\triangle} \underset{\text{lead oxide}}{2PbO} + 4NO_2\uparrow + O_2\uparrow$$

2. A nitrate is heated with concentrated sulphuric acid.

$$\underset{\text{sodium nitrate}}{4NaNO_3} + \underset{\text{sulphuric acid}}{4H_2SO_4} \xrightarrow{\Delta} \underset{\text{sodium hydrogen sulphate}}{4NaHSO_4} + \underset{\text{Water}}{2H_2O} + \underset{\text{nitrogen dioxide}}{4NO_2\uparrow} + \underset{\text{oxygen}}{O_2\uparrow}$$

Tests. (1) Dark brown, heavy gas with a characteristic pungent smell.

(2) Turns a saturated solution of iron (II) sulphate black owing to the formation of black addition product $FeSO_4.NO$.

Experiment 3-4. To test for nitrogen dioxide.

Place a few crystals of lead nitrate in a dry test tube and heat. The crystals decrepitate. The following observations are made :

1. A heavy, dark brown gas is evolved.
2. The gas has a characteristic pungent odour.
3. When the brown gas is passed through a saturated solution of iron (II) sulphate, the light green solution turns black.

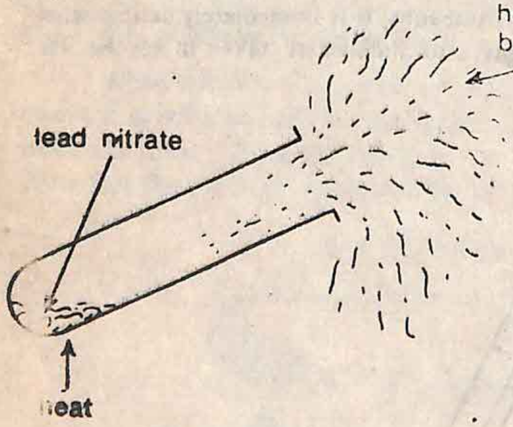

Fig. 3-4. Nitrogen dioxide is evolved when lead nitrate is heated alone.

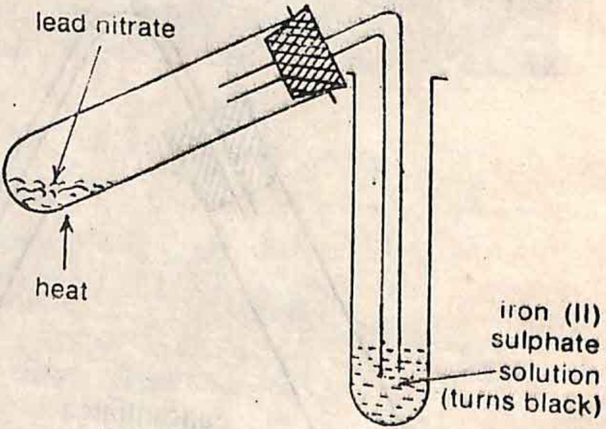

Fig. 3-5. Nitrogen dioxide turns a light green solution of iron (II) sulphate black.

Chlorine, Cl_2

Chlorine is evolved when

1. Concentrated hydrochloric acid is added to potassium permanganate.

$$\underset{\text{potassium permanganate}}{2KMnO_4} + \underset{\text{hydrochloric acid}}{16HCl} \rightarrow \underset{\text{potassium chloride}}{2KCl} + \underset{\text{manganese chloride}}{2MnCl_2} + \underset{\text{water}}{8H_2O} + \underset{\text{chlorine}}{5Cl_2\uparrow}$$

2. Concentrated hydrochloric acid is heated with manganese dioxide, lead dioxide, etc.

$$\underset{\text{manganese dioxide}}{MnO_2} + \underset{\text{hydrochloric acid}}{4HCl} \rightarrow \underset{\text{manganese chloride}}{MnCl_2} + \underset{\text{water}}{2H_2O} + \underset{\text{chlorine}}{Cl_2\uparrow}$$

$$\underset{\text{lead dioxide}}{PbO_2} + \underset{\text{hydrochloric acid}}{4HCl} \rightarrow \underset{\text{lead chloride}}{PbCl_2} + \underset{\text{water}}{2H_2O} + \underset{\text{chlorine}}{Cl_2\uparrow}$$

3. A mixture of a metal chloride and manganese dioxide is heated with concentrated sulphuric acid.

$$\underset{\text{sodium chloride}}{2NaCl} + \underset{\text{manganese dioxide}}{MnO_2} + \underset{\text{sulphuric acid}}{3H_2SO_4} \rightarrow \underset{\text{sodium hydrogen sulphate}}{2NaHSO_4} + \underset{\text{manganese sulphate}}{MnSO_4} + \underset{\text{water}}{2H_2O} + \underset{\text{chlorine}}{Cl_2\uparrow}$$

Tests. 1. Greenish-yellow gas with a characteristic odour and choking effect.

2. Bleaches most litmus paper (red or blue).

3. Turns starch-iodide paper blue. This paper is prepared by dipping filter paper in solutions of potassium iodide and starch. Chlorine displaces iodine from potassium iodide; the iodine produces deep blue colour with starch.

$$\underset{\text{pot. iodide}}{2KI} + \underset{\text{chlorine}}{Cl_2} \rightarrow \underset{\text{pot. chloride}}{2KCl} + \underset{\text{iodine}}{I_2}$$

Experiment 3-5. To test for chlorine.

Place a pinch of manganese dioxide in a dry test tube. Add concentrated hydrochloric acid and holding the test tube by a test tube holder start heating the contents in the burner flame Make the following observations and tests :

1. A greenish-yellow gas is evolved.

2. The gas has a characteristic odour and causes choking.

3. Expose a moist blue litmus paper to the gas; it first turns red and then its colour is discharged.

4. Expose a filter paper strip soaked in a solution containing potassium iodide and starch; a deep blue colour develops on the filter paper strip.

All these tests show that the gas evolved is chlorine.

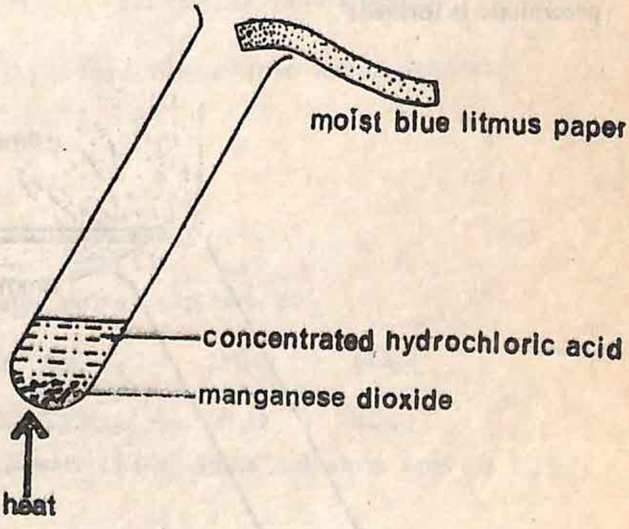

Fig. 3-6. Testing chlorine with moist litmus paper.

6. Hydrogen Chloride, HCl

Hydrogen chloride is formed when a metal chloride is heated with concentrated sulphuric acid.

$$NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl\uparrow$$
$$MgCl_2 + H_2SO_4 \rightarrow MgSO_4 + 2HCl\uparrow$$

Tests. 1. Colourless gas with a pungent odour. Fumes in the air.

2. Turns moist blue litmus paper red.

3. Produces dense white fumes of ammonium chloride when a glass rod which has been dipped in ammonia solution is held in the gas.

$$\underset{\text{ammonia}}{NH_3} + \underset{\text{hydrogen chloride}}{HCl} \longrightarrow \underset{\text{ammonium chloride}}{NH_4Cl}$$

4. Gives a curdy, white precipitate of silver chloride when passed through silver nitrate solution. The precipitate is readily soluble in ammonium hydroxide but does not dissolve in concentrated nitric acid.

$$AgNO_3 + HCl \rightarrow AgCl\downarrow + HNO_3$$

Experiment 3-6. To test for hydrogen chloride.

Place a pinch of sodium chloride in a dry test tube, add a few cm^3 of concentrated sulphuric acid, and heat. A reaction soon starts and white fumes appear at the mouth of the test tube.

1. Observe the colour of the gas inside the test-tube; it is colourless.
2. Carefully smell the gas; it has a sharp odour.
3. Hold a moist blue litmus paper in the gas; it turns red.
4. Dip a glass rod in concentrated ammonia solution and hold the moistened end in the gas; dense white fumes are produced.
5. By means of a delivery tube, pass the gas through silver nitrate solution. A white precipitate is formed.

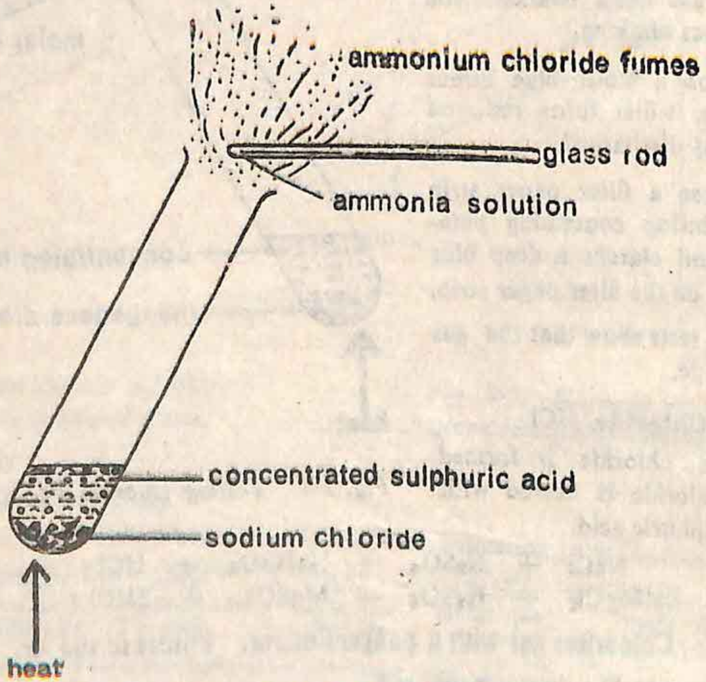

Fig. 3-7. Testing hydrogen chloride with a glass rod which has been dipped in ammonia solution.

Filter the precipitate and divide it in two parts. Treat one part with ammonia solution in a test tube; it readily dissolves. Treat the other part with concentrated nitric acid in another test tube; it does not dissolve.

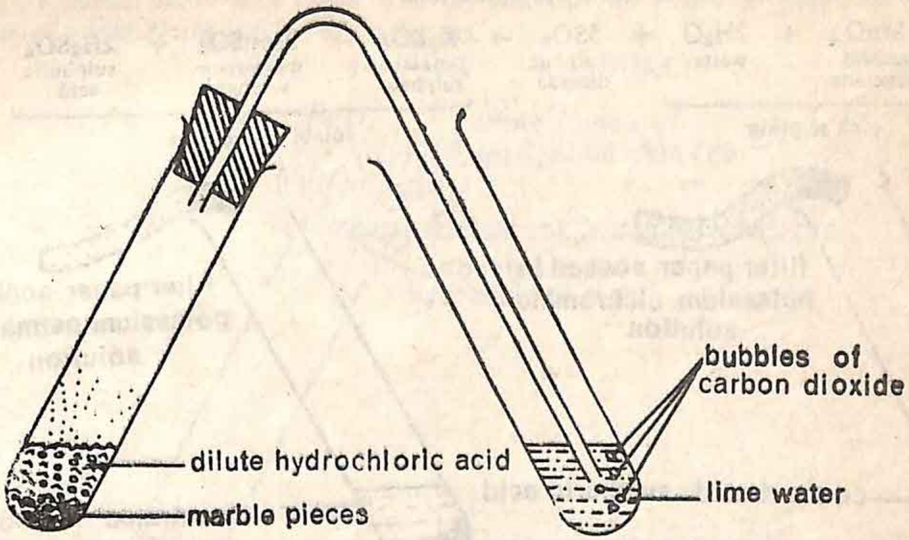

Fig. 3-8. Testing hydrogen chloride by passing it through silver nitrate solution.

7. Sulphur Dioxide, SO_2

Sulphur dioxide is formed when—

1. Sulphur burns in air.

$$S + O_2 \xrightarrow{\Delta} SO_2$$

2. Concentrated sulphuric acid is heated with copper, zinc, etc.

$$Cu + 2H_2SO_4 \xrightarrow{\Delta} CuSO_4 + 2H_2O + SO_2\uparrow$$

$$Zn + 2H_2SO_4 \xrightarrow{\Delta} ZnSO_4 + 2H_2O + SO_2\uparrow$$

3. A sulphite or a hydrogen sulphite is warmed with dilute sulphuric acid or hydrochloric acid.

$$Na_2SO_3 + 2H_2SO_4 \xrightarrow{\Delta} 2NaHSO_4 + H_2O + SO_2\uparrow$$

$$NaHSO_3 + HCl \longrightarrow NaCl + H_2O + SO_2\uparrow$$

Tests. 1. Colourless gas with a sharp and suffocating odour of burning sulphur.

2. Turns moist blue litmus paper blue.

3. Turns filter paper soaked in potassium dichromate solution from orange to green. This is due to the reduction of potassium dichromate (orange colour) to chromic sulphate (green) by sulphur dioxide.

$$\underbrace{\underset{\text{potassium dichromate}}{K_2Cr_2O_7} + \underset{\text{sulphuric acid}}{H_2SO_4}}_{\text{orange solution}} + \underset{\text{sulphur dioxide}}{3SO_2} \rightarrow \underbrace{\underset{\text{potassium sulphate}}{K_2SO_4} + \underset{\text{chromium sulphate}}{Cr_2(SO_4)_3} + H_2O}_{\text{green solution}}$$

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4. Turns filter paper soaked in potassium permanganate solution from purple to colourless. This is due to the reduction of potassium permanganate to potassium sulphate and mangnese sulphate, both of which are colourless.

$$\underbrace{\underset{\text{potassium permaganate}}{2KMnO_4} + \underset{\text{water}}{2H_2O}}_{\text{pink solution}} + \underset{\text{sulphur dioxide}}{5SO_2} \rightarrow \underbrace{\underset{\text{potassium sulphate}}{K_2SO_4} + \underset{\text{manganese sulphate}}{2MnSO_4} + \underset{\text{sulphuric acid}}{2H_2SO_4}}_{\text{colourless solution}}$$

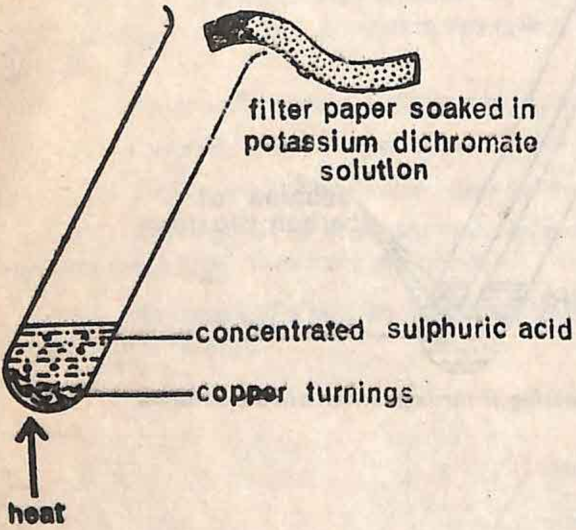
filter paper soaked in potassium dichromate solution
concentrated sulphuric acid
copper turnings
heat

Fig. 3-9. Testing sulphur dioxide with potassium dichromate solution.

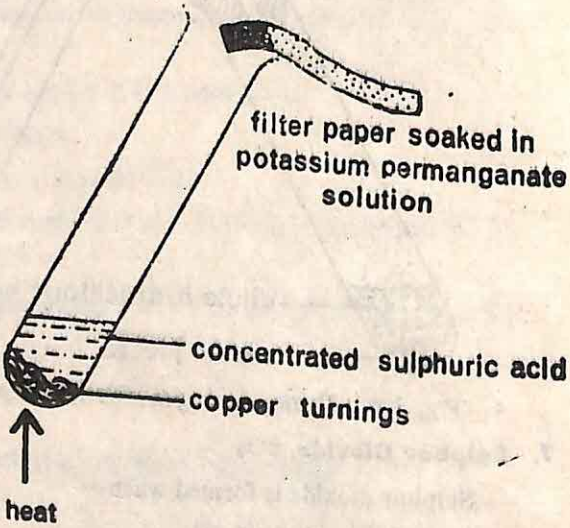
filter paper soaked in potassium permanganate solution
concentrated sulphuric acid
copper turnings
heat

Fig. 3-10. Testing sulphur dioxide with potassium permanganate solution.

Experiment 3-7. To test for sulphur dioxide.

Place some copper turnings in a test tube and add sufficient quantity of concentrated sulphuric acid to cover them. Heat gently in the burner flame. A vigorous reaction sets in and a gas is evolved.

1. Observe the colour of the gas; it is colourless.
2. Carefully smell the gas; it has a sharp, suffocating odour.
3. Dip a strip of filter paper into the potassium dichromate solution. Note that it is stained orange. After allowing it to drain, hold the strip in the gas; the orange colour changes to green.
4. Soak a strip of filter paper in potassium permanganate solution and hold it in the gas; the purple colour is soon discharged.

All these tests show that the gas is sulphur dioxide.

8. Ammonia, NH_3

Ammonia is formed when an ammonium compound is heated with an alkali or slaked lime.

$$NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3\uparrow$$

$$2NH_4Cl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O + 2NH_3\uparrow$$

Tests. 1. Colourless gas with a characteristic pungent odour.

2. Turns moist red litmus paper blue.

3. Produces dense white fumes of ammonium chloride when a glass rod dipped in concentrated hydrochloric acid is held in the gas.

$$NH_3 + HCl \rightarrow NH_4Cl$$

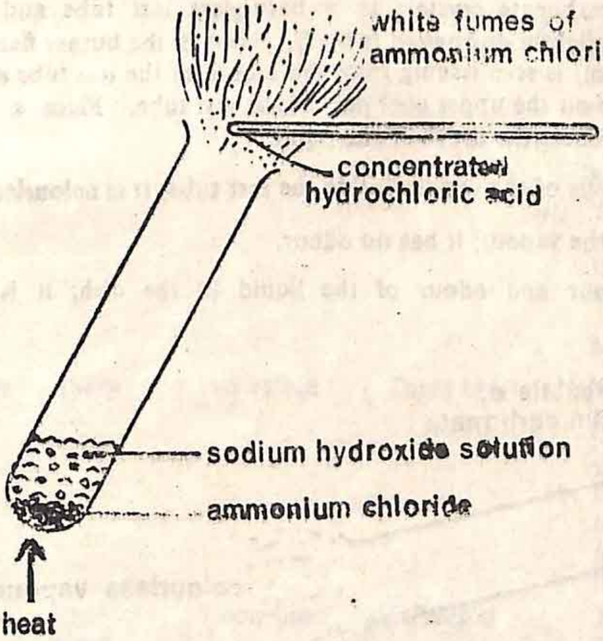
white fumes of ammonium chloride
glass rod
concentrated hydrochloric acid
sodium hydroxide solution
ammonium chloride
heat

Fig. 3-11. Testing ammonia with a glass rod which has been dipped in concentrated hydrochloric acid.

Experiment 3-8. To test for ammonia.

Place some ammonium chloride in a test tube, add a little concentrated sodium hydroxide solution, and heat. Carefully smell by waiving cupped palm over the mouth of the test tube towards your nose. A sharp odour is perceived. This shows that the reaction is occuring with the evolution of ammonia.

1. Observe the colour of the gas; it is colourless.
2. Hold a moist red litmus paper in the gas; it turns blue.
3. Dip a glass rod in concentrated hydrochloric acid, and hold the moistened end in the gas. Dense white fumes of ammonium chloride are formed.

All these tests show that the gas is ammonia,

9. Water Vapour, H_2O

Hydrated salts such as sodium carbonate decahydrate. $Na_2CO_3.10H_2O$, copper sulphate pentahydrate, $CuSO_4$ $5H_2O$, etc. give off the water of hydration as water vapour on heating.

$$Na_2CO_3.10H_2O \xrightarrow{\Delta} Na_2CO_3 + 10\,H_2O\uparrow$$

$$CuSO_4.5H_2O \xrightarrow{\Delta} CuSO_4 + 5H_2O\uparrow$$

Tests 1. Colourless and odourless.

2. Condenses on cool part of the tube as colourless drops. The liquid thus obtained turns white anhydrous copper sulphate blue.

Experiment 3-8. To test for water vapour.

Place some sodium carbonate crystals in a hard-glass test tube and clamp it to a stand so that it is pointing slightly downward (why ?). Heat in the burner flame of a medium length. Soon vapour (steam) is seen issuing from the mouth of the test tube and drops of a colourless liquid are formed on the upper cool part of the test tube. Place a dish under the mouth of the test tube to collect the drops of the liquid.

1. Observe the colour of the vapour inside the test tube; it is colourless*.

2. Carefully smell the vapour; it has no odour.

3. Observe the colour and odour of the liquid in the dish; it is colourless and odourless.

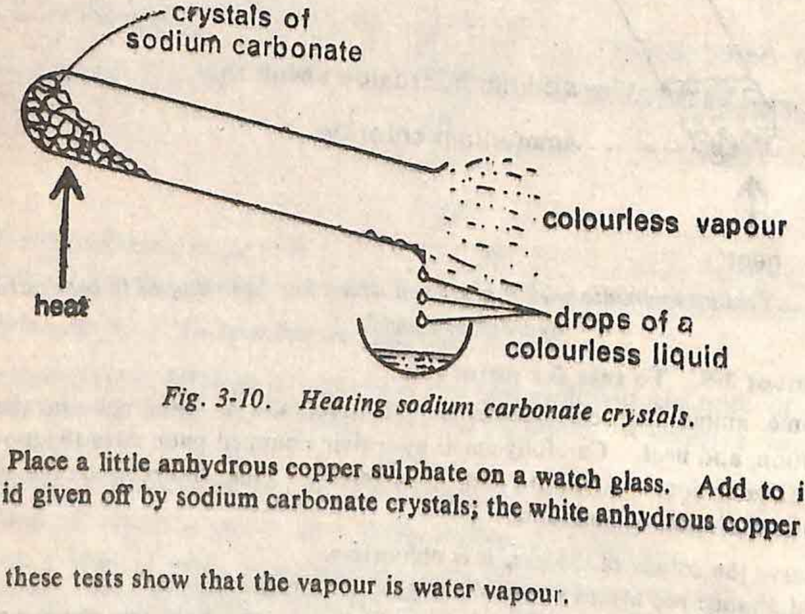

Fig. 3-10. Heating sodium carbonate crystals.

4. Place a little anhydrous copper sulphate on a watch glass. Add to it a few drops of the liquid given off by sodium carbonate crystals; the white anhydrous copper sulphate turns blue.

All these tests show that the vapour is water vapour.

10. Systematic Scheme for the Identification of Some Common Gases

1. Note the colour and the odour of the gas.
2. Test the gas with moist red and blue litmus papers.
3. Apply special test(s).
4. Make deduction.

---

*Water vapour is a colourless gas. In air, it condenses to form very fine droplets of wate[illegible]ich appear like mist. Colourless water vapour is present inside the test tube and between its mouth an[illegible] beginning of the white cloud.

	Colour	*Odour*	*Moist red litmus*	*Moist blue litmus*	*Special test(s)*	*Deduction*
1.	Colourless	Odourless	No action	No action	Introduce lighted splint into the gas.	
					i. The splint is extinguished and the gas burns with a 'pop'.	Hydrogen H_2
					ii. The splint burns more vigorously.	Oxygen O_2
					iii. The splint is extinguished and the gas does not burn. Further confirm by 'lime-water test'	Carbon dioxide CO_2
2.	Colourless	Sharp	No action	Turns red	*i.* Bring a glass rod moistened with NH_4OH; dense, white fumes form. *ii.* Pass the gas into silver nitrate soln.; white ppt. forms.	Hydrogen chloride HCl
3.	Colourless	Characteristic	Turns blue	No action	Bring a glass rod moistened with conc. HCl; dense, white fumes are produced.	Ammonia NH_3
4.	Colourless	Burning sulphur	Bleached	First, turns red, then gets bleached	*i.* Pass the gas into $KMnO_4$ soln.; it is decolorized. *ii.* Pass the gas into acidified pot. dichromate soln.; it turns green.	Sulphur dioxide SO_2
5.	Colourless	Rotten eggs	—	Turns faint red	Pass the gas into lead acetate soln.; black ppt. forms.	Hydrogen sulphide H_2S
6.	Pale-green	Choking	Bleached	Firt turns red, then gets bleached	Pass through silver nitrate soln.; white ppt. forms.	Chlorine Cl_2
7.	Dark-brown	Pungent	No action	Turns red	Pass the gas into $FeSO_4$ soln.; acidified with H_2SO_4; it turns black.	Nitrogen dioxide NO_2

4

Action of Heat on Substances

Heat is an important agent of chemical change. That is to say, heat causes chemical changes to occur in many substances. The most common type of chemical change that heat brings about is the decomposition of a compound into two or more simpler substances. This type of change is called *thermal decomposition*. The oxides of many metals undergo thermal decomposition to give oxygen and either the free metal or a lower oxide of the metal; carbonates of metals (except those of sodium and potassium) yield carbon dioxide and the oxide of the metal; nitrates of heavy metals give off oxygen and nitrogen dioxide leaving behind the metal oxide; hydrated salts lose their water of hydration and change to anhydrous salts; etc. These changes are often accompanied by characteristic colour changes in the substances, and in some cases the solid residue has different colours when hot and when cold.

Some substances do not undergo thermal decomposition but sublime on heating. That is, they change into vapour without melting and the vapour condenses to the same solid substance on the cooler surfaces. Examples of such substances are ammonium chloride, iodine, camphor, etc.

1. Action of Heat on Oxides of Metals

We shall consider here three representative oxides : (1) mercuric oxide, HgO, a bright red powder, (2) red lead, Pb_3O_4, a dark red powder, and (3) lead dioxide, PbO_2.

Experiment 4-1. To study the action of heat on mercuric oxide.

Method. Take about 1 gram of mercuric oxide in a small hard-glass test tube and clamp it to a stand in an almost horizontal position. (Fig. 4-1). Heat the oxide in a small non-luminous Bunsen flame. Take care to see that the test tube is not pointing towards anybody while it is being heated. While heating, introduce a glowing splint half way down the test tube. Should nothing happen immediately, repeat this every few seconds.

After heating, allow the test tube to cool before putting it back in the test-tube rack.

Observations. (*i*) On heating, mercuric oxide gradually darkens in colour, changing from the shiny red colour to almost reddish-black.

(*ii*) The glowing splint introduced into the tube bursts into a flame. This shows that oxygen is being given off.

(*iii*) Silvery droplets of mercury deposit on the upper cool part of the test tube.

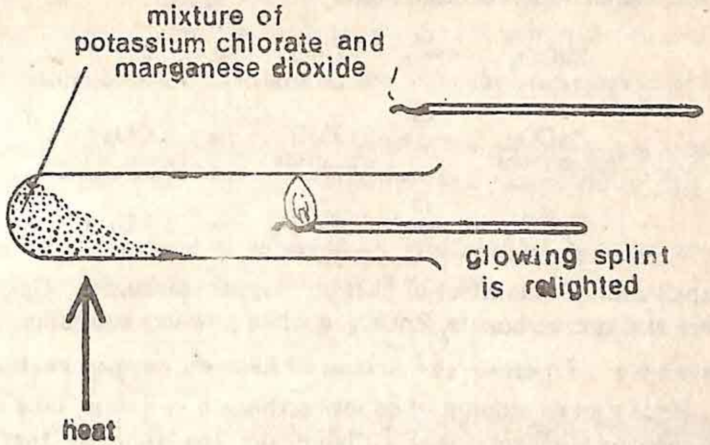

Fig. 4-1. Action of heat on mercuric oxide.

Result. Mercuric oxide breaks up on heating into mercury and oxygen.

$$2HgO \rightarrow 2Hg + O_2\uparrow.$$

Experiment 4-2. To study the action of heat on red lead.

Method. Place a small amout of red lead in small, hard-glass test tube and heat gently as in Experiment 4-1. After heating for some time, put a glowing splint into the tube.

Observations. (*i*) On heating, the colour of red lead changes from red to chocolate brown and finally a yellow-orange powder is formed.

(*ii*) The glowing splint bursts into a flame which shows that the colourless gas evolved is oxygen.

Result. When heated, red lead or lead (II, IV) oxide, Pb_3O_4, breaks up into lead monoxide or lead (II) oxide, PbO, which is a yellow-orange powder, and oxygen.

$$2Pb_3O_4 \xrightarrow{\Delta} 6PbO + O_2\uparrow$$

Experiment 4-3. To study the action of heat on lead dioxide.

Method. Place a small amount of lead dioxide in a test tube and heat it as usual. When the tube is sufficiently hot, put a glowing splint into the test tube.

Observations. (*i*) The chocolate-brown lead dioxide changes to a yellow-orange powder.

(*ii*) The glowing splint rekindles inside the test-tube. This shows that the colourless gas evolved is oxygen.

Result. Lead dioxide or lead (IV) oxide, PbO_2, breaks up on heating to give lead monoxide or lead (II) oxide and oxygen.

$$2PbO_2 \xrightarrow{\Delta} 2PbO + O_2\uparrow$$

2. Action of Heat on Carbonates

Carbonates of common metals, except those of sodium and potassium, decompose on strong heating to give the oxide of the metal and carbon dioxide. The latter can be identified by its action on lime-water which is turned milky.

$$\underset{\text{copper carbonate}}{CuCO_3} \xrightarrow{\Delta} \underset{\text{copper oxide}}{CuO} + \underset{\text{carbon dioxide}}{CO_2\uparrow}$$

$$\underset{\text{zinc carbonat}}{ZnCO_3} \xrightarrow{\Delta} \underset{\text{zinc oxide}}{ZnO} + \underset{\text{carbon dioxide}}{CO_2\uparrow}$$

$$\underset{\text{calcium carbonate}}{CaCO_3} \xrightarrow{\Delta} \underset{\text{calcium oxide}}{CaO} + \underset{\text{carbon dioxide}}{CO_2\uparrow}$$

Here we shall consider the effect of heat on copper carbonate, $CuCO_3$, a light green powdery substance and zinc carbonate, $ZnCO_3$, a white powdery substance.

Experiment 4-4. To study the action of heat on copper carbonate.

Method. Place a small amount of copper carbonate in a test tube and fit it with a one-hole stopper carrying a delivery tube. Clamp the test tube so that its mouth points slightly upwards (Fig. 4-2), and the delivery tube extends to the bottom of a test tube containing 5-6 cm^3 of clear lime-water. Heat the test tube containing copper carbonate, gently at first and then strongly until no further change can be noted.

When the action is complete, allow the tube to cool before putting it back in the test-tube rack and examine the contents.

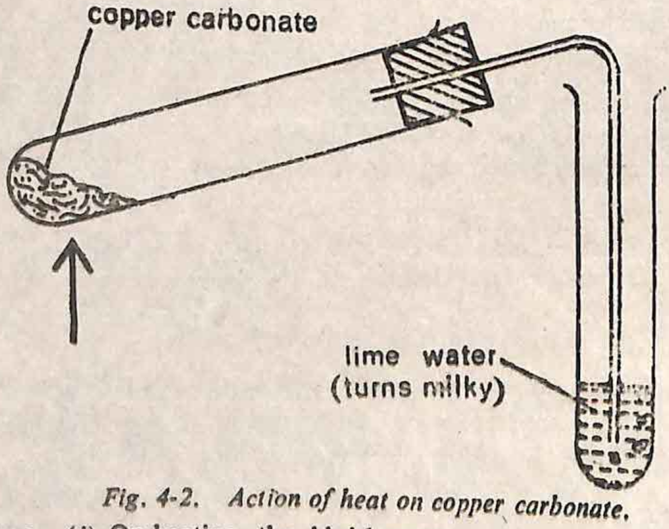

Fig. 4-2. Action of heat on copper carbonate.

Observations. (*i*) On heating, the bluish-green copper carbonate darkens in colour and finally a black powder is left behind as a solid residue.

(*ii*) A colourless gas bubbles through lime water which is turned milky.

Result. On heating, copper carbonate decomposes into copper oxide, a black powder which is left as residue, and carbon dioxide.

$$CuO_3 \xrightarrow{\Delta} CuO + CO_2\uparrow$$

Experiment 4-5. To study the action of heat on zinc carbonate.

Method. Place some zinc carbonate in a test-tube and fit it with a one-hole stopper carrying a delivery tube. Arrange the apparatus as shown in Fig. 3-2. Heat the test-tube containing zinc carbonate in a non-luminous Bunsen flame, gently at first and then strongly.

When the action is complete, allow the tube to cool, and carefully examine the contents.

Observations. (*i*) The residue in the tube is yellow when hot and white when cold.

(*ii*) On heating zinc carbonate, a colourless gas comes off which turns lime-water milky.

Result. When zinc carbonate is heated, it decomposes to give zinc oxide which is yellow when hot and white when cold, and carbon dioxide.

$$ZnCO_3 \xrightarrow{\Delta} ZnO + CO_2\uparrow$$

3. Action of Heat on Hydrates

Crystals of many substances depositing from aqueous solutions contain definite number of water molecules attached to the molecule of the substance that was crystalized. Thus, the blue crystals of copper sulphate contain five molecules of water attached to each molecule of copper sulphate. We indicate this by writing its formula as $CuSO_4.5H_2O$*. The dot in this formula implies a loose attachment of the water molecules. The definite quantity of water contained in the crystals in a state of loose combination is called *water of crystallisation or water of hydration*. A substance that contains water of crystallisation is called a ***hydrated compound*** or a *hydrate*. A few examples of hydrates are :

$CuSO_4.5H_2O$—Copper sulphate *penta*hydrate
$FeSO_4.7H_2O$—Iron (II) sulphate *hepta*hydrate
$Na_2SO_4.10H_2O$—Sodium carbonate *deca*hydrate
$BaCl_2.2H_2O$—Barium chloride *di*hydrate

When hydrated crystals are heated, the water of crystallisation is lost. The substance left behind is said to be *anhydrous*, meaning 'without water'. Hydrates lose their crystalline shape and often their colour, too, on losing their water of crystallisation. For example, when blue crystals of copper sulphate pentahydrate are heated, they lose their water of crystallisation and change to a white, amorphous mass of anhydrous copper sulphate.

$$\underset{\substack{\text{copper sulphate pentahydrate}\\\text{(blue crystals)}}}{CuSO_4.5H_2O} \xrightarrow{\Delta} \underset{\substack{\text{anhydrous copper sulphate}\\\text{(white amorphous mass)}}}{CuSO_4} + \underset{\substack{\text{water of}\\\text{crystallisation}}}{5H_2O\uparrow}$$

All crystals do not contain water of crystallisation. A simple test for its presence is to heat a few crystals in a *dry* test-tube. A film of moisture will appear on the cool portion of the inner wall if the crystalline substance is a hydrate.

We shall now study the action of heat on two representative hydrates : (1) washing soda or sodium carbonate decahydrate, $Na_2CO_3.10H_2O$, a white fine crystalline solid, and (2) copper sulphate crystals or copper sulphate pentahydrate, $CuSO_4.5H_2O$, a blue crystalline solid.

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*All the five molecules of water in $CuSO_4.5H_2O$ are not identically combined; four are attached to the copper (II) ion, Cu^{2+}, and one to the sulphate ion, SO_4^{2-}.

Experiment 4-6. To study the action of heat on copper sulphate crystals.

Method. Place about 5 grams of copper sulphate crystals in a dry test-tube. Clamp the tube so that its mouth is pointing slightly downward. Place a dish below the mouth of the test-tube. Heat the crystals, gently at first and then strongly.

When the action is complete, allow the tube to cool. Observe the residue in the tube. Add a few drops of water to the residue and observe what happens.

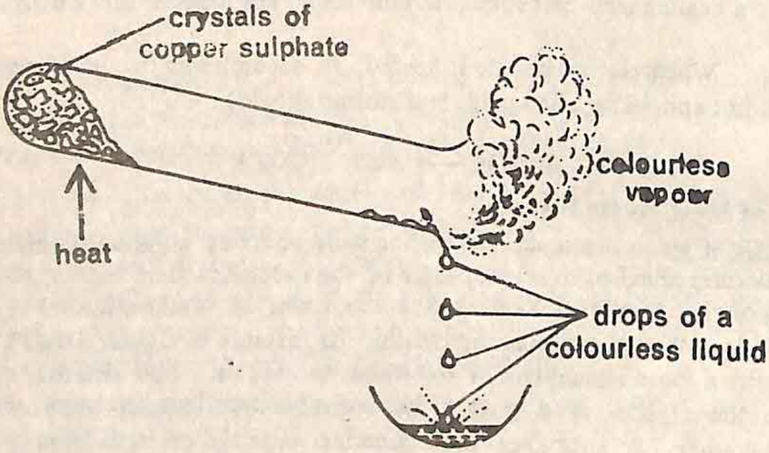

Fig. 4-3. Action of heat on copper sulphate crystals.

Observations. (*i*) When copper sulphate crystals are heated, a colourless and odourless vapour is given off and drops of a colourless liquid appear on the upper cool portion of the test-tube. These drops trickle down and collect in the dish placed below.

(*ii*) The blue crystals of copper sulphate finally fall to a white amorphous mass.

(*iii*) When water is added to the white residue in the test tube, much heat is generated and the colour changes from white to blue. The further addition of water causes the powder to dissolve and form a blue solution.

Results. (*i*) When blue crystals of copper sulphate are heated, water is given off and anhydrous copper sulphate is formed.

$$\underset{\text{(blue crystals)}}{\underset{\text{copper sulphate}}{CuSO_4.5H_2O}} \xrightarrow{\Delta} \underset{\text{(white amorphous powder)}}{\underset{\text{anhydrous copper sulphate}}{CuSO_4}} + 5H_2O \uparrow$$

(*ii*) The white residue of anhydrous copper sulphate on treatment with water turns blue as it changes back to hydrated copper sulphate.

$$\underset{\text{copper sulphate}}{\underset{\text{anhydrous}}{CuSO_4}} + 5H_2O \longrightarrow \underset{\text{sulphate}}{\underset{\text{hydrated copper}}{CuSO_4.5H_2O}}$$

Experiment 4-7. To study the action of heat on washing soda crystals.

Method. Place some washing soda crystals in a dry test tube. Clamp the tube with a slight downward slope towards the mouth. Place a dish below the mouth of the test-tube. Heat the crystals in a non-luminous Bunsen flame, gently at first and then strongly.

When the action is complete, allow the test tube to cool and examine the contents.

Observations. (*i*) On heating, washing soda crystals give off a colourless and odourless vapour.

(*ii*) The vapour partly condenses on the upper cool portion of the tube. The drops of condensed liquid trickle downward and collect in the dish placed below. The liquid is colourless and odourless.

(*iii*) The residue in the tube is white and no longer gives off any vapour on heating.

Results. Washing soda crystals contain water of crystallisation which is given off on heating. The residue left behind is anhydrous sodium carbonate.

$$\underset{\substack{\text{sodium carbonate decahydrate}\\\text{(washing soda crystals)}}}{Na_2CO_3.10H_2O} \xrightarrow{\Delta} \underset{\substack{\text{anhydrous sodium}\\\text{carbonate}}}{Na_2CO_3} + \underset{\substack{\text{water}\\\text{vapour}}}{10H_2O\uparrow}$$

4. Action of Heat on Nitrates

Nitrates of all metals are crystalline solids, readily soluble in water. Some of these contain water of hydration, for example, zinc nitrate, $Zn(NO_3)_2.4H_2O$, while some are anhydrous, for example, lead nitrate, $Pb(NO_3)_2$. All nitrates are decomposed by heat. However, hydrated nitrates first give off their water of hydration before undergoing thermal decomposition.

Nitrates of alkali metals (sodium and potassium) are decomposed by heat into the corresponding nitrites and oxygen.

$$\underset{\text{sodium nitrate}}{2NaNO_3} \xrightarrow{\Delta} \underset{\text{sodium nitrate}}{2NaNO_2} + \underset{\text{oxygen}}{O_2\uparrow}$$

$$\underset{\text{potassium nitrate}}{2KNO_3} \xrightarrow{\Delta} \underset{\text{potassium nitrite}}{2KNO_2} + \underset{\text{oxygen}}{O_2\uparrow}$$

When heated, nitrates, of heavy metals, *e.g.*, those of lead, copper, and zinc, give off a gaseous mixture of oxygen and nitrogen dioxide (a dark-brown gas) leaving behind a solid residue of the corresponding metal-oxide.

$$\underset{\text{lead nitrate}}{2Pb(NO_3)_2} \xrightarrow{\Delta} \underset{\text{lead oxide}}{2PbO} + \underset{\text{nitrogen dioxide}}{4NO_2\uparrow} + \underset{\text{oxygen}}{O_2\uparrow}$$

$$\underset{\text{copper nitrate}}{2Cu(NO_3)_2} \xrightarrow{\Delta} \underset{\text{copper oxide}}{2CuO} + \underset{\text{nitrogen dioxide}}{4NO_2\uparrow} + \underset{\text{oxygen}}{O_2\uparrow}$$

$$\underset{\text{zinc nitrate}}{2Zn(NO_3)_2} \xrightarrow{\Delta} \underset{\text{zinc oxide}}{2ZnO} + \underset{\text{nitrogen dioxide}}{4NO_2\uparrow} + \underset{\text{oxygen}}{O_2\uparrow}$$

Here, we shall study the action of heat on three metal nitrates : (1) zinc nitrate, $Zn(NO_3)_2.4H_2O$, (2) copper nitrate, $Cu(NO_3)_2.3H_2O$, and (3) lead nitrate, $Pb(NO_3)_2$. Of these, zinc nitrate and copper nitrate are crystalline hydrates, while the crystals of lead nitrate are anhydrous. Zinc nitrate and lead nitrate are white crystals while the crystals of copper nitrate are bright, deep blue.

Experiment 4-8. To study the action of heat on zinc nitrate.

Method : Place a small amount of zinc nitrate crystals in a test tube. Clamp the tube so that its mouth is pointing slightly downward. Heat the crystals in a non-luminous flame, gently at first and then strongly.

When the action is complete, allow the tube to cool, and examine the contents.

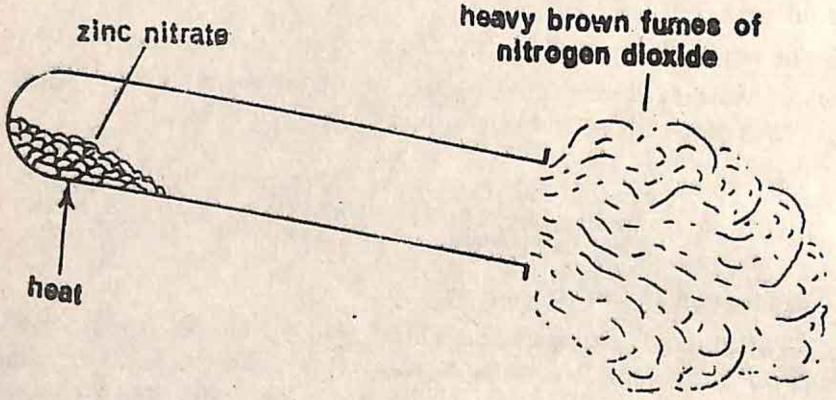

Fig. 4-4. Action of heat on zinc nitrate crystals.

Observations. (*i*) The crystals give off colourless and odourless water vapour which partly condenses on the upper cool portion of the tube.

(*ii*) A heavy, reddish brown gas is given off.

(*iii*) The residue in the tube is yellow when hot and white when cold.

Results. (*i*) Zinc nitrate contains water of crystallisation.

(*ii*) It decomposes on heating to give off nitrogen dioxide, leaving a residue of zinc oxide.

$$\underset{\substack{\text{zinc nitrate}\\\text{(colourless crystals)}}}{2Zn(NO_3)_2} \xrightarrow{\Delta} \underset{\substack{\text{zinc oxide}\\\text{(yellow, when hot;}\\\text{white, when cold)}}}{2ZnO} + \underset{\substack{\text{nitrogen}\\\text{dioxide}}}{4NO_2\uparrow} + \underset{\text{oxygen}}{O_2\uparrow}$$

Experiment 4-9. To study the action of heat on copper nitrate.

Method. Place some crystals of copper nitrate in a test tube. Clamp the tube with a downward slope towards the mouth. Heat the crystals in a non-luminous Bunsen flame, gently at first and then strongly.

When the action is complete, allow the test tube to cool, and examine the contents.

Observations. (*i*) The crystals give off colourless and odourless water vapour on heating. The water vapour partly condenses on the upper cool portion of the tube.

(*ii*) The deep-blue crystals of copper nitrate change to a green, amorphous mass on losing the water vapour.

(*iii*) On further heating, a heavy, dark reddish-brown gas is given off.

(*iv*) The green amorphous mass finally changes to a black mass.

Results. (*i*) Copper nitrate is a hydrated compound, *i.e.*, it contains water of crystallisation.

(*ii*) On heating, copper nitrate first gives off the water of crystallisation and changes to anhydrous copper nitrate which is a green amorphous substance.

$$\underset{\substack{\text{copper nitrate crystals}\\ \text{(blue)}}}{Cu(NO_3)_2.3H_2O} \xrightarrow{\Delta} \underset{\substack{\text{anhydrous copper}\\ \text{nitrate (light green)}}}{Cu(NO_3)_2} + \underset{\substack{\text{water of}\\ \text{crystallization}}}{3H_2O}$$

(*iii*) On continued heating, anhydrous copper nitrate decomposes to give nitrogen dioxide (along with oxygen) and copper (II) oxide (the black residue.)

$$\underset{\substack{\text{copper nitrate}\\ \text{(light green)}}}{2Cu(NO_3)_2} \xrightarrow{\Delta} \underset{\substack{\text{copper (II) oxide}\\ \text{(black)}}}{2CuO} + \underset{\substack{\text{nitrogen}\\ \text{dioxide}}}{4NO_2\uparrow} + \underset{\text{oxygen}}{O_2\uparrow}$$

Experiment 4-10. To study the action of heat on lead nitrate.

Method. Place some lead nitrate crystals in a test tube. Clamp the tube so that it is pointing slightly downwards. Heat the crystals, gently at first and then strongly.

When the action is complete, discontinue heating. Allow the tube to cool, and examine the contents.

Observations. (*i*) On heating, the crystals of lead nitrate break apart with a crackling sound*. However, water is not given off.

(*ii*) A heavy, dark reddish-brown gas is given off on further heating.

(*iii*) The residue in the tube is orange when hot and yellow when cold.

Results. (*i*) Lead nitrate crystals do not contain water of crystallization.

(*ii*) On heating, lead nitrate decomposes to give nitrogen dioxide gas and a yellow-orange solid residue of lead oxide.

$$\underset{\substack{\text{lead nitrate}\\ \text{(white)}}}{2Pb(NO_3)_2} \xrightarrow{\Delta} \underset{\substack{\text{lead oxide}\\ \text{(yellow)}}}{2PbO} + \underset{\substack{\text{nitrogen}\\ \text{dioxide}}}{4NO_2} + \underset{\text{oxygen}}{O_2\uparrow}$$

5. Thermal Dissociation

The breaking up of mercuric oxide into mercury and oxygen by the action of heat is a non-reversible reaction. No amount of cooling or prolonged contact between mercury and oxygen will cause the re-formation of mercuric oxide. Similarly, copper carbonate is split up into copper oxide and carbon dioxide by the action of heat in a non-reversible process. Such reactions are called *thermal decomposition*.

$$2HgO \xrightarrow{\Delta} 2Hg + O_2\uparrow$$

$$2CuCO_3 \xrightarrow{\Delta} 2CuO + 2CO_2\uparrow$$

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*The breaking apart of anhydrous salt on heating with a crackling sound is called *decrepitation*.

When a substance splits up by the action of heat but the products re-form the original substance on cooling, the reaction is called *thermal dissociation.* For example, when ammonium chloride, a white solid, is heated it breaks up into ammonia and hydrogen chloride which recombine on cooling to form ammonium chloride. This is shown by the equation.

$$NH_4Cl \rightleftharpoons NH_3 + HCl$$

The sign $\rightleftharpoons$ in this equation means that the reaction can proceed either way, depending on conditions.

Experiment 4-11. To study the action of heat on ammonium chloride.

Method. Place a very small amount of solid ammonium chloride in a dry test tube. Hold the tube in a slightly inclined position (Fig. 4-5) and heat the solid gently.

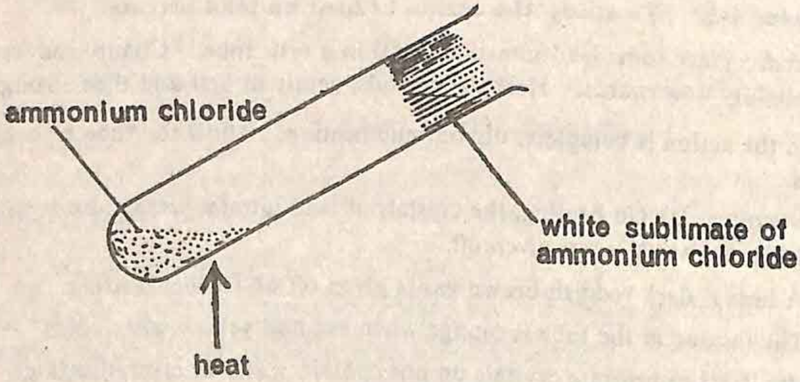

Fig. 4-5. Action of heat on ammonium chloride.

Observations. The solid at the bottom of the tube disappears as it is heated. A white solid appears at the top of the tube, where it is colder. When the condensed solid is reheated, in a test tube, the same thing happens again.

Result. Ammonium chloride breaks up into ammonia and hydrogen chloride under the influence of heat, but the two gases recombine on cooling giving back ammonium chloride as a white sublimate.

$$NH_4Cl \rightleftharpoons NH_3 + HCl$$

Experiment 4-12. To show that ammonium chloride dissociates into ammonia and hydrogen chloride by the action of heat.

Method. Place about 1 g solid ammonium chloride in a dry test tube. Clamp the tube in a nearly horizontal position. Introduce blue and red litmus papers and a loose glass-wool plug as shown in Fig. 4-6. Heat the ammonium chloride gently.

Observations. The ammonium chloride slowly disappears from the bottom of the tube. The neutral litmus paper goes red while the red litmus paper goes blue.

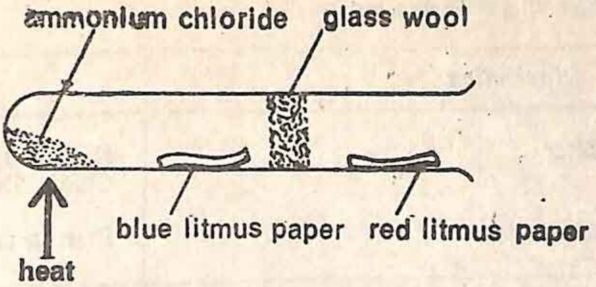

Fig. 4-6. Dissociation af ammonium chloride by heat.

Result. When heated, ammonium chloride splits up into ammonia (which turns red litmus paper blue) and hydrogen chloride (which turns blue litmus paper red).

6. Action of Heat on an Unknown Solid Substance

Useful information leading to the identification of an unknown substance can often be obtained by studying the action of heat upon it.

Place some of the unknown substance in a small hard-glass test tube. Holding the tube nearly horizontal, heat it gently over a non-luminous burner flame. Shake gently to ensure uniform heating of the solid. Observe the effects, if any, of gentle heating. Then heat the sample more strongly and again make your observations.

When the change appears to be complete, allow the tube and its contents to cool. Note whether the residue remains the same as when hot, or changes in some other way.

Use a fresh tube for each substance under investigation.

Record your observations with respect to the following changes, and draw inferences from the observations made :

I. *Condensed water, decrepitation, sublimation and melting*

Observation	*Inference*
1. Evolution of colourless vapour which condenses higher up the tube as colourless droplets of a liquid. The liquid turns anhydrous copper sulphate blue showing that it is water.	A hydrated salt, *e.g.*, $Na_2CO_3.10H_2O$, $CuSO_4.5H_2O$, $FeSO_4.7H_2O$, etc.
2. The solid decrepitates (*i.e.*, breaks into smaller pieces with crackling sound).	Lead nitrate or sodium chloride.
3. The solid sublimes, *i.e.*, changes to vapour which condenses higher up the tube as solid.	
(*a*) White sublimate.	Ammonium chloride
(*b*) Grey sublimate (from violet vapour).	Iodine
(*c*) Grey sublimate which, when rubbed, changes to shiny globules.	Mercury compound.
4. The solid melts.	A nitrate or a halide.

II. *Colour change on gentle heating.*

Observation	*Inference*
1. Blue to dirty white	Hydrated copper sulphate, $CuSO_4.5H_2O$
2. Light green to dirty white	Ferrous sulphate, $FeSO_4.7H_2O$

III. *Colour of residue (hot and cold) left on strong heating.*

Observation	*Inference*
1. Yellow when hot, white when cold.	Zinc compound, *e.g.*, $ZnCO_3$, $Zn(NO_3)_2$, etc.
2. Orange when hot, yellow when cold.	Lead compound, *e.g.*, $Pb(NO_3)_2$
3. Black (hot and cold).	Copper compound, *e.g.*, $Cu(NO_3)_2$
4. Black when hot, rust red when cold.	Iron compound, *e.g.*, $FeSO_4$

IV. *Gas evolved (Note colour and odour and apply special tests)*

Observation	*Inference*
1. A colourless, odourless gas.	
(*a*) Rekindles a glowing splint (O_2)	Mercuric oxide (if red) or a nitrate (if colourless)
(*b*) Turns lime water milky (CO_2)	Metal carbonate (except that of sodium or potassium)
2. A colourless gas with characteristic smell.	
(*a*) Pungent, turns red litmus blue (NH_3)	Ammonium salt, *e.g.*, NH_4Cl
(*b*) Pungent, turns yellow potassium dichromate green (SO_2)	Sulphite, *e.g.*, $ZnSO_3$ or sulphate,, *e.g.*, $FeSO_4$
(*c*) Bad eggs, turns lead acetate black (H_2S)	Sulphide
(*d*) Very sharp ; turns blue litmus red and produces white fumes with NH_4OH (HCl)	Hydrated chloride
3. Dark brown, heavy gas (NO_2)	Nitrate

5

Action of Sulphuric Acid on Substances

Sulphuric acid dissolves the oxides of most metals forming their soluble sulphates. Thus, it dissolves copper (II) oxide which is changed to soluble copper (II) sulphate ($CuO+H_2SO_4 \rightarrow CuSO_4+H_2O$). Dioxides and peroxides of metals dissolve in concentrated sulphuric acid on heating forming their sulphates with the liberation of oxygen ($2MnO_2+2H_2SO_4 \overset{\Delta}{\rightarrow} 2MnSO_4+2H_2O+O_2$). Most salts of metals are decomposed by sulphuric acid with evolution of gases that are characteristic of the acid radical or the anion in the salt. Identification of these gases, therefore, leads to the deduction of the acid radical in the salt.

Dilute sulphuric acid decomposes the salts of unstable, volatile acids such as the carbonates, sulphides, and sulphites liberating carbon dioxide, hydrogen sulphide, and sulphur dioxide respectively. Carbonates are decomposed in the cold; the sulphides and sulphites are decomposed on warming.

Such salts as the chlorides and nitrates are not decomposed by dilute sulphuric acid even on heating. However, these are decomposed when heated with concentrated sulphuric acid; chlorides give off pale-green chlorine gas and the nitrates give off heavy, dark brown fumes of nitrogen dioxide.

Dilute sulphuric acid reacts with active metals such as zinc, iron and magnesium in cold. In the reaction, the metal dissolves as its sulphate and hydrogen is evolved with brisk effervescence.

1. Action of Dilute Sulphuric Acid on Carbonates

When a carbonate is treated with dilute sulphuric acid, carbon dioxide is liberated with sudden effervescence or 'fizz'. As the reaction proceeds, the carbonate dissolves forming the sulphate*. If the acid is in excess, the carbonate will completely dissolve to form a clear solution, and then there will be no more effervescence. For example,

$$\underset{\text{sodium carbonate}}{Na_2CO_3} + H_2SO_4 \rightarrow \underset{\text{sodium sulphate}}{Na_2SO_4} + H_2O \quad CO_2\uparrow$$

$$\underset{\text{copper carbonate}}{CuCO_3} + H_2SO_4 \rightarrow \underset{\text{copper sulphate}}{CuSO_4} + H_2O + CO_2\uparrow$$

$$\underset{\text{zinc carbonate}}{ZnCO_3} + H_2SO_4 \rightarrow \underset{\text{zinc sulphate}}{ZnSO_4} + H_2O + CO_2\uparrow$$

Carbon dioxide is identified by the following observations : (1) it is a colourless and odourless gas, (2) it extinguishes a lighted splint, and (3) it turns lime water milky.

Experiment 5-1. To study the action of dilute sulphuric acid on copper carbonate.

Procedure. Take a small quantity of copper carbonate (about 0·5 gram) in a test tube and add a little dilute**, sulphuric acid (about 4-5 cm^3).

Observations. There is a sudden effervescence on adding dilute sulphuric acid and small bubbles of a colourless gas are formed. The gas has no odour.

When a lighted splint is introduced in the gas, it is extinguished.

When the gas is passed through lime-water, it turns milky.

Deductions. The tests show that the colourless and odourless gas evolved is carbon dioxide. Hence, the given substance is a carbonate.

Repeat the experiment with sodium carbonate and zinc carbonate in separate test tubes.

2. Action of dilute sulphuric acid on sulphides

Dilute sulphuric acid liberates hydrogen sulphide from metal sulphides. The reaction often takes place in cold but warming may be necessary in some cases. As in the case of

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*In the case of lead carbonate and calcium carbonate, there is a sudden evolution of carbon dioxide when dilute sulphuric acid is added, but the reaction soon stops. The reason is that the lead sulphate and calcium sulphate formed are insoluble. These sulphates cover the original carbonates and prevent the action of the acid on them.

$$\underset{\text{lead carbonate}}{PbCO_3} + H_2SO_4 \rightarrow \underset{\substack{\text{lead sulphate}\\ \text{(insoluble)}}}{PbSO_4\downarrow} + H_2O + CO_2\uparrow$$

$$\underset{\text{calcium carbonate}}{CaCO_3} + H_2SO_4 \rightarrow \underset{\substack{\text{calcium sulphate}\\ \text{(insoluble)}}}{CaSO_4} + H_2O + CO_2\uparrow$$

However, if dilute nitric acid is used in place of dilute sulphuric acid, whole of these carbonates dissolve with the liberation of carbon dioxides as their nitrates produced are soluble.

$$\underset{\text{lead carbonate}}{PbCO_3} + 2HNO_3 \rightarrow \underset{\substack{\text{lead nitrate}\\ \text{(soluble)}}}{Pb(NO_3)_2} + H_2O + CO_2\uparrow$$

$$\underset{\text{calcium carbonate}}{CaCO_3} + 2HNO_3 \rightarrow \underset{\substack{\text{calcium nitrate}\\ \text{(soluble)}}}{Ca(NO_3)_2} + H_2O + CO_2\uparrow$$

**The sulphuric acid used is 1M strong

carbonates, the metal sulphides dissolove in sulphuric acid forming soluble sulphates. For example,

$$\underset{\text{zinc sulphide}}{ZnS} + H_2SO_4 \rightarrow \underset{\text{zinc sulphate}}{ZnSO_4} + H_2S\uparrow$$

$$\underset{\text{iron sulphide}}{FeS} + H_2SO_4 \rightarrow \underset{\text{iron(II) sulphate}}{FeSO_4} + H_2S\uparrow$$

Hydrogen sulphide is identified by the following observations : (1) it is a colourless gas having the unpleasant smell of 'rotten eggs' (2) it produces a black stain* on a strip of filter paper moistened with lead acetate solution.

Experiment 5-2. To study the action of dilute sulphuric acid on iron sulphide.

Procedure. Place a small piece of iron sulphide in a test tube and add a little dilute sulphuric acid. Warm gently as the reaction is rather slow in cold.

Observations. (1) A colourless gas is evolved with mild effervescence.

(2) The gas has the smell of 'rotten eggs'.

(3) When a filter paper strip moistened with lead acetate solution is exposed to the gas, a black stain is produced on it.

(4) Iron sulphide dissolves in dilute sulphuric acid to form a light green solution.

Deductions. Iron sulphide reacts with dilute sulphuric acid to give off hydrogen sulphide and dissolves forming a light green solution of iron (II) sulphate.

$$FeS + H_2SO_4 \rightarrow FeSO_4 + H_2S\uparrow$$

Repeat the experiment with zinc sulphide.

3. Action of Dilute Sulphuric Acid on Sulphites

When a sulphite, *e.g.*, sodium sulphate, is warmed with dilute sulphuric acid, sulphur dioxide gas is evolved and the sulphite is changed to sulphate. For example,

$$\underset{\text{sodium sulphite}}{Na_2SO_3} + H_2SO_4 \rightarrow \underset{\text{sodium sulphate}}{Na_2SO_4} + H_2O + SO_2\uparrow$$

$$\underset{\text{zinc sulphite}}{ZnSO_3} + H_2SO_4 \rightarrow \underset{\text{zinc sulphate}}{ZnSO_4} + H_2O + SO_2\uparrow$$

Sulphur dioxide is identified by the following observations : (1) it is a colourless gas, having the choking smell of 'burning sulphur', (2) it produces a green stain on a filter paper soaked in potassium dichromate solution, and (3) it decolorizes the pink solution of potassium permanganate. (For explanations of these tests and chemical equations see Chapter 3).

Experiment 5-3. To study the action of dilute sulphuric acid on sodium sulphite.

Procedure. Place a small quantity of sodium sulphite in a clean, dry test tube. Add a little dilute sulphuric acid. Since no reaction occurs in cold, warm the mixture.

---

*The black stain is due to lead sulphide formed by the action of hydrogen sulphide on lead acetate.

$$\underset{\text{lead acetate}}{(CH_3.COO)_2Pb} + \underset{\substack{\text{hydrogen}\\\text{sulphide}}}{H_2S} \rightarrow \underset{\substack{\text{lead sulphide}\\\text{(black stain)}}}{PbS} + \underset{\text{acetic acid}}{2CH_3.COOH}$$

Observations. (1) On warming sodium sulphite with dilute sulphuric acid a colourless gas is evolved.

(2) The gas has the pungent and choking smell of 'burning sulphur'.

(3) When a filter paper strip soaked in potassium dichromate solution is exposed to the gas, it turns green.

(4) When the gas is passed through potassium permanganate solution, its pink colou is discharged.

(5) Sodium sulphite dissolves in the acid to form a colourless solution.

Deductions. When sodium sulphite is heated with dilute sulphuric acid, sulphur dioxide is evolved and sodium sulphite dissolves forming a colourless solution of sodium sulphate.

$$Na_2SO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O + SO_2\uparrow$$

Repeat the experiment with zinc sulphite.

4. Action of Dilute Sulphuric Acid on Metals

Dilute sulphuric acid reacts with active metals (*i.e.*, metals above hydrogen in the activity series), such as magnesium, zinc, iron, etc., in the cold. In these reactions, hydrogen is evolved with brisk effervescence and the metal passes into solution as its soluble sulphate. For example,

$$\underset{\text{magnesium}}{Mg} + H_2SO_4 \rightarrow \underset{\text{magnesium sulphate}}{MgSO_4} + H_2\uparrow$$

$$\underset{\text{zinc}}{Zn} + H_2SO_4 \rightarrow \underset{\text{zinc sulphate}}{ZnSO_4} + H_2\uparrow$$

$$\underset{\text{iron}}{Fe} + H_2SO_4 \rightarrow \underset{\text{iron (II) sulphate}}{FeSO_4} + H_2\uparrow$$

Hydrogen is identified by the following observations : (1) it is a colourless and odourless gas, and (2) when a lighted splint is brought near it, the splint is extinguished but the gas burns with a 'pop'.

Experiment 5-4. To study the action of dilute sulphuric acid on iron.

Procedure. Place a little iron filings (about 0·5 g) in a test tube and add a little dilute sulphuric acid.

Observations. (1) Brisk effervescence immediately ensues and tiny bubbles of a colourless gas are formed.

(2) The gas has no odour.

(3) When a lighted splint is brought near the mouth of the test tube, the splint is extinguished but the gas burns with a 'pop'. The flame of the burning gas is almost invisible and pale-blue.

(4) Iron dissolves in the acid forming a light green solution of iron (II) sulphate.

Deductions. Iron reacts with dilute sulphuric acid to liberate hydrogen and dissolves in it forming iron (II) sulphate.

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$$

5. Action of Concentrated Sulphuric Acid on Chlorides

Concentrated sulphuric acid reacts with metal chlorides on heating to liberate hydrogen chloride and convert the metal chlorides into corresponding sulphates. For example,

$$\underset{\text{calcium chloride}}{CaCl_2} + H_2SO_4 \xrightarrow{\Delta} \underset{\text{calcium sulphate}}{CaSO_4} + 2HCl\uparrow$$

$$\underset{\text{zinc chloride}}{ZnCl_2} + H_2SO_4 \xrightarrow{\Delta} \underset{\text{zinc sulphate}}{ZnSO_4} + 2HCl\uparrow$$

$$\underset{\text{ammonium chloride}}{2NH_4Cl} + H_2SO_4 \xrightarrow{\Delta} \underset{\text{ammonium sulphate}}{(NH_4)_2SO_4} + 2HCl\uparrow$$

Hydrogen chloride is identified by the follow observations : (1) it is a colourless gas with a pungent smell, (2) it produces dense white fumes of ammonium chloride in contact with a glass rod moistened with ammonia solution, and (3) when passed into silver nitrate solution, it produces a curdy, white precipitate of silver chloride. (See Chapter 3, for explanations and chemical equations).

Experiment 5-5. To study the action of concentrated sulphuric acid on sodium chloride.

Procedure. Take a small quantity of sodium chloride (0·4-0·5 g) in a test tube and add a little concentrated sulphuric acid (4-5 cm^3). Heat the mixture gently, but do not boil.

Observations. (1) A colourless gas is evolved when sodium chloride is heated with concentrated sulphuric acid.

(2) The gas has a pungent smell.

(3) When a glass rod dipped in ammonia solution is brought near the mouth of the test tube, dense white fumes are formed.

(4) When the gas is passed through silver nitrate solution, a curdy, white precipitate is formed.

Deduction. When sodium chloride is heated with concentrated sulphuric acid, hydrogen chloride is evolved.

$$NaCl + H_2SO_4 \xrightarrow{\Delta} NaHSO_4 + HCl\uparrow$$

Repeat the experiment with zinc chloride.

6. Action of Concentrated Sulphuric Acid on Nitrates

When a nitrate is heated with concentrated sulphuric acid, nitric acid is liberated. The nitric acid is partially decomposed by heat into water, oxygen, and nitrogen dioxide. The gaseous mixture evolved appears as heavy, dark brown fumes due to the nitrogen dioxide, NO_2.

For example,

$$NaNO_3 + H_2SO_4 \xrightarrow{\Delta} NaHSO_4 + HNO_3$$

$$4HNO_3 \xrightarrow{\Delta} \underbrace{2H_2O\uparrow + O_2\uparrow + 4NO_2\uparrow}_{\text{heavy, dark brown fumes}}$$

If copper turnings are added to the hot mixture of concentrated sulphuric acid and the nitrate, the brown fumes becomes copious as more nitrogen dioxide is formed by the reaction of copper and nitric acid.

$$Cu + 4HNO_3 \xrightarrow{\Delta} Cu(NO_3)_2 + 2H_2O + 2NO_2\uparrow$$

The heavy brown fumes have a characteristic, pungent smell.

Experiment 5-6. To study the action of concentrated sulphuric acid on sodium nitrate.

Procedure. Place about 0·4-0·5 g of sodium nitrate in a clean, dry test tube. Add a little concentrated sulphuric acid, and heat.

Observations. (1) Heavy, brown fumes are evolved.

(2) The fumes have a characteristic pungent smell.

(3) On adding copper turnings to the hot reaction mixture, the brown fumes become more copious.

Deduction. On heating with concentrated sulphuric acid, sodium nitrate gives off brown fumes of nitrogen dioxide.

$$4NaNO_3 + 4H_2SO_4 \xrightarrow{\Delta} 4NaHSO_4 + 2H_2O + O_2\uparrow + 4NO_2\uparrow$$

Repeat the exercise with copper nitrate and zinc nitrate.

6

Flame Tests

Flame tests are based on the fact that certain salts, particularly the chlorides, when heated directly in a non-luminous flame, volatilise and ionise, and their cations thus produced impart characteristic colours to the flame. Thus, these tests provide a simple method of identifying such cations.

The best way of putting the salt in the flame is to carry it on a platinum (or nichrome) wire. A short length, about 4 cm, is sealed into the end of a glass rod which serves as the handle. The reason of using platinum is that it is very non-reactive and, therefore, neither gets oxidized in the flame, nor does it impart any colour to the flame itself.

1. Procedure of Flame Test

A flame test is performed in the following steps :

(*i*) *Cleansing the platinum wire.* Take out a small quantity of concentrated hydrochloric acid on a *clean* watch glass. Dip the free end of the platinum wire in the acid and

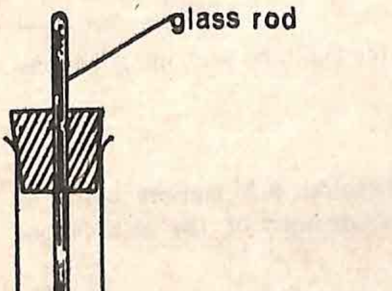

Fig. 6-1. The platinum wire is kept immersed in a little concentrated hydrochloric acid in a test tube so that it will remain clean

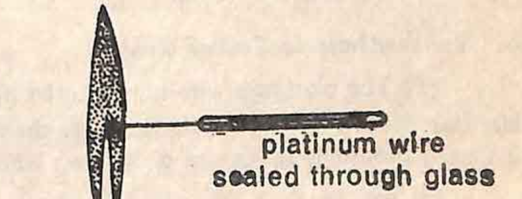

Fig. 6-2. Introducing the paste of the substance and concentrated hydrochloric acid on the end of platinum wire in the flame.

introduce it into a non-luminous Bunsen flame. Dip again into the hydrochloric acid, and re-heat. Repeat until no colour is produced in the flame.

(*ii*) *Introducing the substance in the flame.* Place a very small quantity of the substance (a few specks would be sufficient) on a *clean* watch glass. Moisten the substance with a few drops of concentrated hydrochloric acid and mix thoroughly with a glass rod to make a paste. Pick a little of the paste on the end of the cleansed platinum wire. Hold the end in the bottom side of the *non-luminous* flame (Fig. 6-2). Observe the colour imparted to the flame first with naked eyes and then through a blue glass.

2. Deduction from Flame Coloration

Make appropriate deduction from your observation of the flame coloration according to the table given below :

Flame coloration		*Metal/cation indicated*
With naked eyes	*Through blue glass*	
1. Bright golden yellow, very persistent	Nil	Sodium, Na$^+$
2. Violet	Pink	Potassium, K$^+$
3. Orange-red; flickering	Very light green	Calcium, Ca^{2+}

After performing the test, replace the platinum wire in the test tube containing concentrated hydrochloric acid (Fig. 6-1).

3. Precautions in Flame Tests

(1) The platinum wire must not be dipped into the hydrochloric acid reagent bottle as any impurities on it would contaminate the entire acid. A small quantity of the acid (about 2·3 cm^3) should be taken out on a clean watch glass.

(2) The acid used in cleaning platinum wire must not be used for making the paste of the substance for the flame test; it has been contaminated with traces of salts on the wire. Throw away this acid, thoroughly wash the watch glass, and pour out a fresh lot of the acid for the test.

(3) The flame test must not be applied with a match stick, splint, etc. If a platinum wire is not available, a 'nichrome' (an alloy) wire or the lead of a pencil can be used as a substitute.

(4) The golden-yellow flame of sodium is very persistent. A yellow flame appearing momentarily should be neglected; it may be due to impurities in the given substance.

(5) If the given substance is coloured, do not apply flame rest.* Why ?

---

*All common compounds of sodium, potassium, and calcium are either colourless or white.

4. A Sample Experiment

Experiment 6-1. To identify the metals in samples A, B and C by flame tests.

Method. Clean the platinum wire by dipping it into concentrated hydrochloric acid taken out on a clean watch glass and heating in a non-luminous flame. Repeat until colour no longer appears in the flame.

Make a paste of a small quantity of the sample A with concentrated hydrochloric acid on a clean watch glass. Pick up a little of the paste on the end of the wire and hold it in the bottom side of the non-luminous flame.

Repeat with samples B and C.

Observations. Sample A imparts a violet colour to the flame which appears pink through a blue glass.

Sample B imparts a golden yellow colour to the flame. The colour is masked when viewed through a blue glass.

Sample C imparts an orange-red colour to the flame which appears light green through a blue glass.

Deductions. (1) Sample A contains potassium.

(2) Sample B contains sodium.

(3) Sample C contains calcium.

7

Action of Sodium Hydroxide Solution on Substances

Sodium hydroxide, NaOH, is an important alkali. It is readily soluble in water, and dissociates in solution into Na^+ and OH^- ions ($NaOH \rightleftharpoons Na^+ + OH^-$). When its solution is added to the solutions of salts containing Ca^{2+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, Pb^{2+}, Zn^{2+}, etc., the OH^- ions combine with these metal ions to produce insoluble metal-hydroxides which are precipitated. These metal-hydroxides can be easily distinguished from one another by their colours and by their being soluble or insoluble in excess of sodium hydroxide solution. Thus, these metal ions can be easily identified by observing the action of sodium hydroxide on their solutions.

Ammonium salts (NH_4^+), when heated with sodium hydroxide solution, give off ammonia. As such, an ammonium salt can be easily identified by heating it with sodium hydroxide solution, and delecting the ammonia evolved by suitable tests.

1. General Procedure

Make a solution of the given substance in water or, if insoluble in water, in dilute hydrochloric acid. Take a few cm^3 of the solution in a test tube. Add sodium hydroxide solution* in small portions, shake the test tube after each addition. If a precipitate forms, note its colour and texture, and then continue adding sodium hydroxide solution until the volume of the mixture has trebled. The mixture now contains an excess of sodium hydroxide. Note whether or not the precipitate dissolves.

If a precipitate does not form on addition of the sodium hydroxide solution, heat the mixture. Observe by smell if ammonia is evolved or not.

2. Action on Calcium Salts

When sodium hydroxide solution is added to the solution of a calcium salt, a white ecipitate of calcium hydroxide is formed. The precipitate is thin and consists of fine particles. or example,

$$\underset{\substack{\text{calcium}\\ \text{chloride}}}{CaCl_2} + \underset{\substack{\text{sodium}\\ \text{hydroxide}}}{2NaOH} \rightarrow \underset{\substack{\text{calcium}\\ \text{hydroxide}}}{Ca(OH)_2 \downarrow} + \underset{\substack{\text{sodium}\\ \text{chloride}}}{2NaCl}$$

---

*A 2M sodium hydroxide solution is used.

In general,

$$Ca^{2+} + 2OH^- \rightarrow Ca(OH)_2\downarrow$$

The precipitate of calcium hydroxide does not dissolve in the excess of sodium hydroxide solution.

Experiment 7-1. To study the effect of sodium hydroxide on calcium salts.

Method. Take a pinch of calcium chloride in a test tube. Dissolve it in the least possible quantity of water to obtain a saturated solution. To the solution add sodium hydroxide in small amounts at a time till it is an excess.

Observation. A white precipitate consisting of very fine particles is formed. The precipitate does not dissolve in the excess of the alkali.

Inference. The Ca^{2+} ions form a white precipitate of $Ca(OH)_2$ on adding alkali.

Repeat the experiment with calcium nitrate. Similar result will be obtained.

3. Action on Copper Salts

On adding sodium hydroxide solution to the solution of a copper salt, *i.e.*, a solution containing Cu^{2+} ions, a blue precipitate of copper (II) hydroxide, $Cu(OH)_2$, is formed. For example,

$$\underset{\substack{\text{copper (II)}\\ \text{sulphate}}}{CuSO_4} + \underset{\substack{\text{sodium}\\ \text{hydroxide}}}{2NaOH} \rightarrow \underset{\substack{\text{copper (II)}\\ \text{hydroxide}}}{Cu(OH)_2\downarrow} + Na_2SO_4$$

In general,

$$Cu^{2+} + 2OH^- \rightarrow Cu(OH)_2\downarrow$$

The precipitate of copper (II) hydroxide does not dissolve in the excess of the sodium hydroxide solution. However, if the mixture containing the blue precipitate is boiled, it turns black. This is because copper hydroxide changes into black copper (II) oxide by the action of heat.

$$\underset{\substack{\text{copper (II) hydroxide}\\ \text{(blue ppt.)}}}{Cu(OH)_2} \xrightarrow{\Delta} \underset{\substack{\text{copper (II) oxide}\\ \text{(black ppt.)}}}{CuO} + H_2O$$

Experiment 7-2. To study the action of sodium hydroxide on copper salts.

Method. Dissolve about 0·3-0·4 g of powdered copper sulphate in minimum quantity of water in a test tube. To the solution add sodium hydroxide solution, a few drops at a time and with shaking, till in excess.

Observation. A light blue precipitate is formed that does not dissolve in the excess of the alkali. When the reaction mixture is heated, it turns black.

Inference. The Cu^{2+} ions form a precipitate of $Cu(OH)_2$ on the addition of alkali. On heating, it changes to black copper (II) oxide, CuO.

Repeat the experiment with copper chloride and copper nitrate. Identical Result will be obtained in each case.

4. Action on Iron (II) Salts

Addition of sodium hydroxide solution to the solution of a ferrous or iron (II) salt, produces a light green precipitate of iron (II) hydroxide, $Fe(OH)_2$. The texture of the precipitate is gelatinous.

$$\underset{\text{iron (II) sulphate}}{FeSO_4} + \underset{\text{sodium hydroxide}}{2NaOH} \rightarrow \underset{\text{iron (II) hydroxide}}{Fe(OH)_2\downarrow} + \underset{\text{sodium sulphate}}{Na_2SO_4}$$

In general,

$$Fe^{2+} + 2OH^- \rightarrow Fe(OH)_2\downarrow$$

The precipitate of iron (II) hydroxide does not dissolve in the excess of sodium hydroxide solution.

Experiment 7-3. To observe the action of sodium hydroxide solution on iron (II) salts.

Method. Dissolve a pinch of ferrous sulphate, or iron (II) sulphate, in a small quantity of distilled water in a clean test tube. To the light green solution formed add sodium hydroxide solution, in small lots and with shaking after each addition, till the alkali is in excess.

Observation. A dirty green, gelatinous precipitate, is formed. The precipitate does not dissolve in the excess of sodium hydroxide solution.

Inference. Fe^{2+} ions react with sodium hydroxide solution to produce a dirty green precipitate of $Fe(OH)_2$.

Repeat the experiment with iron (II) chloride. Similar result will be obtained.

5. Action on Iron (III) Salts

When sodium hydroxide solution is added to the solution of an iron (III) salt (or, a ferric salt) containing the Fe^{3+} ions, a brown precipitate of iron (III) hydroxide or ferric hydroxide is formed. For example,

$$\underset{\text{iron (III) chloride}}{FeCl_3} + \underset{\text{sodium hydroxide}}{3NaOH} \rightarrow \underset{\text{iron (III) hydroxide}}{Fe(OH)_3\downarrow} + \underset{\text{sodium chloride}}{3NaCl}$$

In general,

$$Fe^{3+} + 3OH^- \rightarrow Fe(OH)_3\downarrow$$

The precipitate of iron (III) hydroxide does not dissolve in the excess of sodium hydroxide solution.

Experiment 7-4. To study the action of sodium hydroxide solution on iron (III) or ferric salts.

Method. Dissolve a few crystals of ferric chloride or iron (III) chloride in a small quantity of water in a clean test tube. To the yellow solution obtained add sodium hydroxide solution, a few drops at a time and with shaking, until it is in excess.

Observation. A brown precipitate is formed that does not dissolve in the excess of the alkali.

Inference. Fe^{3+} ions form a brown precipitate of $Fe(OH)_3$ when alkali is added.

Repeat the experiment with iron (III) nitrate. Identical result will be obtained.

6. Action on Lead Salts

On adding sodium hydroxide to the solution a lead salt, *i.e.*, a solution containing Pb^{2+} ions, a white precipitate of lead hydroxide is formed. For example,

$$\underset{\text{lead nitrate}}{Pb(NO_3)_2} + \underset{\text{sodium hydroxide}}{2NaOH} \rightarrow \underset{\text{lead hydroxide}}{Pb(OH)_2\downarrow} + \underset{\text{sodium nitrate}}{2NaNO_3}$$

In general, $Pb^{2+} + 2OH^{+} \rightarrow Pb(OH)_2\downarrow$

The white precipitate of lead hydroxide dissolves in sodium hydroxide solution added in excess. This is because it reacts with sodium hydroxide to form soluble lead plumbite.

$$\underset{\substack{\text{lead hydroxide}\\ \text{(insoluble)}}}{Pb(OH)_2} + \underset{\substack{\text{sodium}\\ \text{hydroxide}}}{2NaOH} \rightarrow \underset{\substack{\text{lead plumbite}\\ \text{(soluble)}}}{Na_2PbO_2} + \underset{\text{water}}{2H_2O}$$

Experiment 7-5. To study the action of sodium hydroxide solution on lead salts.

Method. Dissolve a few crystals of lead nitrate in a small quantity of distilled water. To the solution add sodium hydroxide solution, a few drops at a time and with shaking after each addition, till the alkali is in excess.

Observation. A white precipitate is first formed which dissolves in the excess of the alkali yielding a clear solution.

Inference. Pb^{2+} ions produce a white precipitate of $Pb(OH)_2$ on treatment with sodium hydroxide solution. The precipitate dissolves in the excess of the alkali.

7. Action on Zinc Salts

When sodium hydroxide solution is added to the solution of a zinc salt, *i.e.*, a solution containing Zn^{2+} ions, a white precipitate of zinc hydroxide, $Zn(OH)_2$, is first formed. For example,

$$\underset{\substack{\text{zinc}\\ \text{chloride}}}{ZnCl_2} + \underset{\substack{\text{sodium}\\ \text{hydroxide}}}{2NaOH} \rightarrow \underset{\substack{\text{zinc}\\ \text{hydroxide}}}{Zn(OH)_2\downarrow} + \underset{\substack{\text{sodium}\\ \text{chloride}}}{2NaCl}$$

In general,

$$Zn^{2+} + 2OH^{-} \rightarrow Zn(OH)_2\downarrow$$

On adding sodium hydroxide solution in excess, the precipitate of zinc hydroxide dissolves owing to the formation of soluble sodium zincate.

$$\underset{\substack{\text{zinc hydroxide}\\ \text{(insoluble)}}}{Zn(OH)_2} + \underset{\substack{\text{sodium}\\ \text{hydroxide}}}{2NaOH} \rightarrow \underset{\substack{\text{sodium zincate}\\ \text{(soluble)}}}{Na_2ZnO_2} + \underset{\text{water}}{2H_2O}$$

Thus, the observations on the addition of sodium hydroxide solution are identical for both lead and zinc salts—first, formation of a white precipitate and then dissolution of the precipitate in the excess of the alkali. However, distinction between the two salts can be easily made by tests given in Chapter 9.

Experiment 7·6. To study the action of sodium hydroxide solution on zinc salts.

Method. Dissolve a pinch of zinc sulphate in a small quantity of water in a clean test tube. To the solution obtained add sodium hydroxide solution, drop by drop and with shaking, till it is in excess.

Observation. A white precipitate is formed that dissolves in the excess of sodium hydroxide solution.

Inference. On adding sodium hydroxide to a solution containing Zn^{2+} ions a white precipitate of $Zn(OH)_2$ is formed which dissolves when more alkali is added.

$$ZnSO_4 + 2NaOH \rightarrow Zn(OH)_2\downarrow + Na_2SO_4$$
$$Zn(OH)_2 + 2NaOH \rightarrow Na_2ZnO_2 + 2H_2O$$

Repeat the experiment with zinc chloride and zinc nitrate; identical results will be obtained.

8. Action on Ammonium Salts

When an ammonium salt or its solution in water is heated with sodium hydroxide solution, ammonia is evolved. For example,

$$\underset{\text{ammonium chloride}}{NH_4Cl} + \underset{\text{sodium hydroxide}}{NaOH} \xrightarrow{\Delta} \underset{\text{sodium chloride}}{NaCl} + \underset{\text{water}}{H_2O} + \underset{\text{ammonia}}{NH_3\uparrow}$$

In general,

$$\underset{\text{(from the ammonium salt)}}{NH_4^+} + \underset{\text{(from the alkali)}}{OH^-} \xrightarrow{\Delta} H_2O + NH_3\uparrow$$

The ammonia evolved is readily recognized by its characteristic pungent smell. Further confirmation can be made by exposing moist blue litmus paper to the gas which is turned blue, and by holding a glass rod dipped in concentrated hydrochloric acid in the gas when dense, white fumes of ammonium chloride are formed (See Chapter 3).

Experiment 7-7. To study the action of sodium hydroxide solution on ammonium salts.

Method. Take a pinch of ammonium sulphate in a clean test tube. Add a few drops of sodium hydroxide solution and heat gently. Carefully smell the gas evolved.

Observation. A colourless gas having pungent smell is evolved. The gas turns moist red litmus paper blue and produces dense, white fumes in contact with a glass rod moistened with concentrated sulphuric acid. These observations show that the gas evolved is ammonia.

$$(NH_4)_2SO_4 + 2NaOH \xrightarrow{\Delta} Na_2SO_4 + 2H_2O + 2NH_3$$
$$NH_3 + HCl \longrightarrow NH_4Cl \text{ (white fumes)}$$

Inference. When heated with sodium hydroxide solution, ammonium salts give off ammonia.

Repeat the experiment with ammonium chloride. Similar result will be obtained.

8

Qualitative Analysis

The identification of an unknown substance is a very important branch of chemistry which plays an important role in the manufacture of countless products by chemical processes. **Qualitative analysis**, as it is called, has been brought to a high level of perfection, and highly sensitive tests have been devised to detect substances even when they are present in traces.

The qualitative analysis of a salt involves the identification of its metallic part or the basic radical and the non-metallic part or the acid radical. If a salt is regarded as the product of neutralization of an acid and a base; the metallic part of the salt is provided by the base and is, therefore, called the *basic radical*, and the non-metallic part by the acid which is, therefore, called the *acid radical*. For example, the salt sodium chloride is formed by the neutralization of sodium hydroxide (the base) and hydrochloric acid.

$$\underset{\substack{\text{hydrochloric}\\ \text{acid}}}{HCl} + \underset{\substack{\text{sodium hydroxide}\\ \text{base}}}{NaOH} \longrightarrow \underset{\substack{\text{sodium chloride}\\ \text{(salt)}}}{NaCl} + \underset{\text{water}}{H_2O}$$

The sodium in the salt comes from the base and is the basic radical; chlorine comes from the acid, and is the acid radical, called the chloride radical. In solution, the basic and acid radicals exist as the corresponding ions, *e.g.*, Na^+ and Cl^-. Some common basic and acid radicals are :

Acid radicals	*Basic radicals*
1. Carbonate, CO_3^{2-}	1. Calcium, Ca^{2+}
2. Sulphide, S^{2-}	2. Copper, Cu^{2+}
3. Sulphite, SO_3^{2-}	3. Iron (II), Fe^{2+}
4. Chloride, Cl^-	4. Iron (III), Fe^{3+}
5. Nitrate, NO_3^-	5. Lead, Pb^{2+}
6. Sulphate, SO_4^{2-}	6. Zinc, Zn^{2+}
	7. Ammonium, NH_4^+

1. Qualitative Analysis of Salts

The qualitative analysis of an unknown salt involves the identification of its constituent acid and basic radicals by applying suitable tests. Some of the usual tests are—

1. Action of dilute sulphuric acid or hydrochloric acid.
2. Action of concentrated sulphuric acid.
3. Tests with special reagents in solution.

} for acid radicals

4. Action of sodium hydroxide solution.
5. Action of ammonium hydroxide.
6. Flame tests (Chapter 6).

} for basic redicals

Before proceeding to tests for the acid and basic radicals in the salt, note its colour and whether it feels heavy or light.

(1) If the salt is blue, copper (Cu^{2+}) is indicated.
(2) If the given salt is light green, iron (II) is indicated.
(3) If the given salt is yellow, iron (III) is indicated.
(4) If the given salt feels exceptionally heavy, lead is indicated.
(5) If the salt feels exceptionally light, it may be magnesium carbonate.

2. Dilute Sulphuric Acid Test

The test is applied by treating the given salt with dilute sulphuric acid, first in the cold and then by using heat. The identity of the acid radical is determined by identifying the gaseous product of the reaction, if any, by suitable tests.

Procedure. Take a small quantity (about 0·2 g) of the given salt in a dry test tube and add 3-4 cm^3 of dilute sulphuric acid. First observe the change, if any, in the cold, and apply suitable tests to identify the gas evolved. If there is no change in the cold, warm (*do not boil*) the contents of the test tube, and observe if any gas is evolved. Identify the gas by applying suitable tests.

Deductions. On the basis of the observations it is possible to identify the gas and therefrom to deduce the acid radical in the salt.

Table 8-1. *Dilute Sulphuric acid test*

Observation	*Deduction of acid radical*	*Confirmatory test*
1. Sudden effervescence ensues in cold, and a *colourless* and odourless gas is evolved.	Carbonate CO_3^{2-}	Pass the gas through lime water; it is turned milky.
2. A *colourless gas with the smell of rotten eggs* is evolved on warming.	Sulphide S^{2-}	Soak a strip of filter paper in lead acetate solution. Hold it over the mouth of the test tube; a black stain is produced on the paper.
3. *A colourless gas with the smell of burning sulphur* (or burning matchstick) is evolved on warming.	Sulphite SO_3^{2-}	(*i*) Hold a moist blue litmus paper near the mouth of the test tube; it turns red. (*ii*) Soak a strip of filter paper in potassium dichromate solution. Hold it over the mouth of the test tube; the moist part becomes green.

Note. Dilute hydrochloric acid can be used in place of dilute sulphuric acid; the same gases will be evolved.

Explanations. 1. The carbonates give off carbon dioxide (CO_2) by the action of dilute sulphuric acid in the cold. Carbon dioxide is a colourless and odourless gas. For example,

$$ZnCO_3 + H_2SO_4 \rightarrow ZnSO_4 + H_2O + CO_2\uparrow$$

Carbon dioxide turns lime-water (calcium hydroxide) milky as it reacts with it to form insoluble, white particles of calcium carbonate.

$$\underset{\substack{\text{calcium hydroxide}\\ \text{(lime water)}}}{Ca(OH)_2} + \underset{\substack{\text{carbon}\\ \text{dioxide}}}{CO_2} \rightarrow \underbrace{\underset{\text{calcium carbonate}}{CaCO_3\downarrow} + \underset{\text{water}}{H_2O}}_{\text{milky liquid}}$$

2. When a sulphide is warmed with dilute sulphuric acid, hydrogen sulphide (H_2S) is evolved. It is a colourless gas with a characteristic smell of rotten eggs. For example,

$$FeS + H_2SO_4 \xrightarrow{\Delta} FeSO_4 + H_2S\uparrow$$

Hydrogen sulphide blackens lead acetate paper due to the formation of black lead sulphide.

$$\underset{\text{lead acetate}}{(CH_3.COO)_2Pb} + \underset{\substack{\text{hydrogen}\\ \text{sulphide}}}{H_2S} \longrightarrow \underset{\substack{\text{lead sulphide}\\ \text{(black)}}}{PbS} + \underset{\text{acetic acid}}{2CH_3.COOH}$$

3. When a sulphite is warmed with sulphuric acid, sulphur dioxide gas is evolved. It is a colourless gas with a sharp, pungent, and suffocating smell of burning sulphur. For example,

$$\underset{\text{sodium sulphite}}{Na_2SO_3} + H_2SO_4 \xrightarrow{\Delta} \underset{\text{sodium sulphate}}{Na_2SO_4} + H_2O + \underset{\text{sulphur dioxide}}{SO_2\uparrow}$$

Sulphur dioxide combines with water to form sulphurous acid. This is why it turns moist blue litmus paper red.

$$H_2O + SO_2 \rightarrow H_2SO_3 \text{ (sulphurous acid)}$$

Sulphur dioxide turns filter paper soaked in potassium dichromate solution (orange colour) green by reducing it to chromic sulphate.

$$\underbrace{\underset{\substack{\text{potassium}\\ \text{dichromate}}}{K_2Cr_2O_7} + \underset{\substack{\text{sulphuric}\\ \text{acid}}}{H_2SO_4}}_{\text{orange solution}} + \underset{\substack{\text{sulphur}\\ \text{dioxide}}}{3SO_2} \rightarrow \underbrace{\underset{\substack{\text{potassium}\\ \text{sulphate}}}{K_2SO_4} + \underset{\substack{\text{chromium}\\ \text{sulphate}}}{Cr_2(SO_4)_3} + H_2O}_{\text{green solution}}$$

2-1. **Another confirmatory test for sulphite.** Take about 1 cm^3 of dilute sulphuric acid in a clean test tube. Add to it 4-5 drops of potassium permanganate solution. A pink solution will result. Add a pinch of the sulphite salt, *e.g.*, sodium sulphite, and gently warm. The pink colour of the solution will be discharged, *i.e.*, the solution will become colourless.

In this test sulphur dioxide is evolved by the reaction between the sulphite and hot, dilute sulphuric acid. Sulphur dioxide is a reducing reagent. It reduces potassium permanganate to colourless manganese sulphate and, therefore, the solution becomes colourless.

$$\underset{\text{sodium sulphite}}{Na_2SO_3} + 2H_2SO_4 \xrightarrow{\Delta} 2NaHSO_4 + H_2O + \underset{\text{sulphur dioxide}}{SO_2\uparrow}$$

$$\underbrace{\underset{\substack{\text{potassium}\\ \text{permanganate}}}{2KMnO_4} + 2H_2O}_{\text{pink solution}} + 5SO_2 \rightarrow \underbrace{\underset{\substack{\text{potassium}\\ \text{sulphate}}}{K_2SO_4} + \underset{\substack{\text{manganese}\\ \text{sulphate}}}{2MnSO_4} + \underset{\substack{\text{sulphuric}\\ \text{acid}}}{2H_2SO_4}}_{\text{colourless solution}}$$

Experiment 8-1. To find the gases evolved when the given substance A1 is treated with dilute sulphuric acid and make deductions from the observation.

Method. Place a portion (about 0·2-0·3 g) of the substance A1 in a clean, dry test tube. Add about 4 cm³ of dilute sulphuric acid to the substance in the test tube.

Observation. A brisk effervescence occurs in the cold due to the evolution of a colourless and odourless gas.

Identification of the gas evolved. Pass the gas through lime water by means of a delivery tube. The lime-water turns milky.

This test shows that the gas evolved is *carbon dioxide.*

Deduction. The acid radical in the given substance A1 is carbonate, CO_3^{2-}.

Experiment 8-2. To find the gas evolved when the given substances A2 is treated with dilute sulphuric acid and make deductions from the observations.

Method. Place about 0·2—0·3 g of the substance A2 in a clean, dry test tube. Add about 4-5 cm³ of dilute sulphuric acid. Warm gently, if no reaction occurs in the cold.

Observation. No reaction appears to occur in the cold. On warming, a colourless gas is evolved with the smell of rotten eggs.

Identification of the gas evolved. Soak a strip of filter paper in lead acetate solution, and hold it over the month of the test tube. A black stain appears on the strip of filter paper.

This test shows that the gas evolved is *hydrogen sulphide.*

Deduction. The acid radical in the given substance is sulphide, S^{2-}.

Experiment 8-3. To identify the acid radical present in the given substance A3 by the dilute sulphuric acid test.

Method. Take pinch of the substance A3 in a clean, dry test tube. Add about 2 cm³ of dilute sulphuric acid. Warm gently, as no reaction occurs in the cold.

Observation. On warming, a colourless gas with the smell of burning sulphur is evolved.

Identification of the gas evolved. (1) Hold a moist blue litmus paper near the mouth of the test tube; it turns red.

(2) Hold a filter paper strip soaked in potassium dichromate solution near the mouth of the test tube; it turns green.

These tests show that the gas evolved is sulphur dioxide.

Deduction. The acid radical in substance A3 is sulphite, SO_3^{2-}.

3. Concentrated Sulphuric Acid Test

As described in Chapter 5, this test is used for the identification of the acid radicals chloride and sulphate. When heated with concentrated sulphuric acid, chlorides give off hydrogen chloride and the nitrates give off nitrogen dioxide. These gases can be identified by suitable tests. Obviously, if a substance gives off hydrogen chloride when heated with concentrated sulphuric acid, it is a chloride, and if it gives off nitrogen dioxide, it is a nitrate.

Procedure. Place a portion of the given substance (about 0·2—0·3 g) in a clean, dry test tube in a holder. Add 3-4 cm^3 of concentrated sulphuric acid to the substance, and heat *gently*. *Do not boil the acid* and be careful to keep the test tube pointed away from everyone.

Apply suitable tests to identify the gas evolved.

Deductions. Make deductions pertaining to concentrated sulphuric acid test according to the Table 8-2.

Table 8-2. *Concentrated sulphuric acid test*

	Observation	*Acid radical indicated*	*Confirmatory tests*
1.	A *colourless gas with a pungent smell* is evolved.	Chloride Cl^-	(*i*) Hold a moist blue litmus paper near the mouth of the test tube. The litmus paper turns red. (*ii*) Dip glass rod in ammonia solution and hold the mositened end near the mouth of the test tube. Dense white fumes are produced.
2.	*Dense, brown fumes with a characteristic pungent smell* are evolved.	Nitrate NO_3^-	(*i*) Add copper turnings to the hot reaction mixture in the test tube. The evolution of fumes becomes more copious. (*ii*) Soak a filter paper in freshly prepared saturated solution of iron (II) sulphate in water. Black colour is produced on the filter paper.

Explanations. 1. Hydrogen chloride is an acidic gas and, therefore, it turns moist blue litmus paper red.

When hydrogen chloride comes in contact with ammonia, dense, white fumes of ammonium chloride are formed.

$$\underset{\text{ammonia}}{NH_3} + \underset{\text{hydrogen chloride}}{HCl} \longrightarrow \underset{\text{ammonium chloride}}{NH_4Cl}$$

2. When a nitrate is heated with concentrated sulphuric acid, nitric acid is liberated. The nitric acid is partially decomposed by heat into water, oxygen, and nitrogen dioxide.

$$\underset{\text{zinc nitrate}}{Zn(NO_3)_2} + H_2SO_4 \longrightarrow ZnSO_4 + \underset{\text{nitric acid}}{2HNO_3}$$

$$4HNO_3 \xrightarrow{\Delta} 2H_2O + O_2\uparrow + 4NO_2\uparrow$$

The brown colour of the gas evolved is due to nitrogen dioxide, NO_2.

On adding copper turnings to the hot reaction mixture of a nitrate and concentrated sulphuric acid, the copper and the free nitric acid interact to give copious fumes of nitrogen dioxide.

$$\underset{\text{copper}}{Cu} + \underset{\substack{\text{(from nitrate and}\\ \text{sulphuric acid)}}}{4HNO_3} \rightarrow \underset{\text{copper nitrate}}{Cu(NO_3)_2} + \underset{\text{water}}{2H_2O} + \underset{\substack{\text{nitrogen}\\ \text{dioxide}}}{2NO_2\uparrow}$$

Nitrogen dioxide is reduced to nitric oxide, NO, by iron (II) sulphate. The nitric oxide combines with unreacted iron (II) sulphate to form a black coloured addition compound, $FeSO_4.NO$.

3-1. Other confirmatory tests for chlorides. (1) Mix about 0·2 g of the chloride with about an equal amount of manganese dioxide. Take the mixture in a dry test tube and add a little concentrated sulphuric acid. Heat the mixture. Chlorine gas will evolve. It is readily recognized by its pale-green colour and unpleasant, suffocating odour. For example, if the chloride is sodium chloride—

$$2NaCl + MnO_2 + 3H_2SO_4 \xrightarrow{\Delta} 2NaHSO_4 + MnSO_4 + 2H_2O + Cl_2\uparrow$$

Following tests may be employed for confirming the identity of chlorine gas :

1. The gas turns a moist blue litmus paper first red and then bleaches it colourless.

The reason is that chlorine reacts with water to form hydrochloric and hypochlorous acids ($Cl_2 + H_2O \rightarrow HCl + HClO$). These acids turn blue litmus red. Hypochlorous acid, $HClO$, soon decomposes to gives nascent oxygen ($HClO \rightarrow HCl + O$) which bleaches the litmus paper.

2. Dip a filter paper strip in a mixed solution of starch and potassium iodide. The paper thus obtained is called 'starch-iodide' paper. Hold the starch-iodide paper near the mouth of the test tube. A blue colour will appear.

The explanation is that chlorine displaces iodine from potassium iodide solution. The displaced iodine produces a dark blue colour with starch.

$$\underset{\text{pot. iodide}}{2KI} + \underset{\text{chlorine}}{Cl_2} \rightarrow \underset{\text{pot. chloride}}{2KCl} + \underset{\text{iodine}}{I_2}$$

2. Dissolve a pinch of the given salt* in distilled water. To the solution add silver nitrate solution. A curdy, white precipitate of silver chloride will form. For example,

$$\underset{\text{sodium chloride}}{NaCl} + \underset{\text{silver nitrate}}{AgNO_3} \rightarrow \underset{\text{sodium nitrate}}{NaNO_3} + \underset{\text{er chloride}}{AgCl\downarrow}$$

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*The chlorides given to you will all be soluble in water.

Divide the precipitate in two parts. To the first part add concentrated nitric acid; the precipitate will not dissolve. To the other part add ammonium hydroxide; the precipitate will readily dissolve.

8-2. 'Ring test' for nitrate. Dissolve a pinch of the given nitrate in distilled water in a test tube. Add nearly an equal volume of freshly prepared saturated solution of iron (II) sulphate. Hold the test tube at an angle (Fig. 8-1) and carefully pour concentrated sulphuric acid down the side of the test tube. A brown ring will appear at the junction of the two liquids.

The dark brown colour of the 'ring' is due to the addition compound $FeSO_4.NO$.

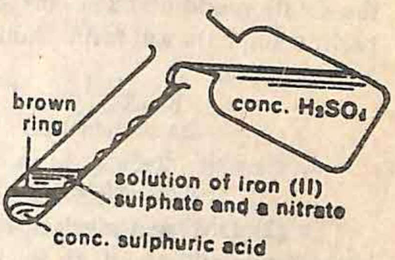

Fig. 8-1. 'Ring test' for nitrates.

Experiment 8-4. To find the gases evolved when the given substance A reacts with concentrated sulphuric acid and make deductions from the observations.

Method. Take a small portion (about 0·2—0·3 g) of the given substance in a clean dry test tube in a holder. Add about 3-4 cm^3 of concentrated sulphuric acid, and heat gently.

Observation. A colourless gas with pungent smell is evolved.

Identification of the gaseous product. (1) Hold moist blue litmus paper over the mouth of the test tube; it goes red.

(2) Bring a glass rod moistened with ammonia solution near the mouth of the test tube; white fumes are produced.

(3) Pass the gas through delivery tube into silver nitrate solution; a white precipitate is formed.

Treat the precipitate with ammonia solution; the precipitate readily dissolves,

All these tests show that the gas evolved is hydrogen chloride.

Deduction. The acid radical in the given substance A is chloride.

Experiment 8-5. To find the gas evolved when the given substance B is heated with concentrated sulphuric acid and make deductions from the observations.

Method. Take a small amount of the given substance in a clean, dry test tube. Add about 3-4 cm^3 of concentrated sulphuric acid, and heat gently.

Observation. A brown gas with a characteristic pungent smell is evolved.

Identification of the gas. 1. Pass the gas through delivery tube into a freshly prepared saturated solution of iron (II) sulphate; the solution turns dark black.

2. Add some copper turnings to the hot reaction mixture in the test tube; the brown fumes become copious.

These tests show that the gas evolved is nitrogen dioxide.

Deduction. The acid radical in the given substance B is nitrate, NO_3^-.

4. Test for Sulphate Radical

(1) Dissolve a pinch of the given sulphate* in 2 cm^3 of distilled water in a test tube. To the solution add about 1 cm^3 of barium chloride solution. A thick, white precipitate of barium sulphate will form Add dilute hydrochloric acid to the precipitate; it will not dissolve. For example,

$$\underset{\text{sodium sulphide}}{Na_2SO_4} + \underset{\text{barium chloride}}{BaCl_2} \rightarrow \underset{\text{sodium chloride}}{2NaCl} + \underset{\text{barium sulphate}}{BaSO_4\downarrow}$$

$$\underset{\text{zinc sulphate}}{ZnSO_4} + \underset{\text{barium chloride}}{BaCl_2} \rightarrow \underset{\text{zinc chloride}}{ZnCl_2} + \underset{\text{barium sulphate}}{BaSO_4\downarrow}$$

(2) Disolve a pinch of the given sulphate in 2 cm^3 of distilled water in a test tube. To the solution add about an equal volume of lead nitrate solution in distilled water. A thick, white precipitate of lead sulphate will form. Add concentrated nitric acid to the precipitate; it will not dissolve. For example,

$$\underset{\text{sodium sulphate}}{Na_2SO_4} + \underset{\text{lead nitrate}}{Pb(NO_3)_2} \rightarrow \underset{\text{sodium nitrate}}{2NaNO_3} + \underset{\text{lead sulphate}}{PbSO_4\downarrow}$$

$$\underset{\text{zinc sulphate}}{ZnSO_4} + \underset{\text{lead nitrate}}{Pb(NO_3)_2} \rightarrow \underset{\text{zinc nitrate}}{Zn(NO_2)_2} + \underset{\text{lead sulphate}}{PbSO_4\downarrow}$$

Experiment 8-6. To test whether the given salt A7 contains sulphate radical or not.

Method. Dissolve a small amount (about 0·5 g) of the salt A7 in 5 cm^3 of distilled water. Divide the solution in two parts.

i. To one part of the solution add about an equal volume of barium chloride solution followed by dilute hydrochloric acid.

ii. To the other part add about an equal volume of lead nitrate solution followed by concentrated nitric acid.

Observations. (*i*) A white precipitate is formed on adding barium chloride solution. The precipitate does not dissolve in hydrochloric acid.

(*ii*) A white precipitate is formed on adding lead nitrate solution. The precipitate does not dissolve in nitric acid.

Deduction. The salt A7 contains sulphate radical, SO_4^{2-}.

5. Identification of Bssic Radicals

A simple method for the identification of the basic radicals Ca^{2+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, Pb^{2+}, and Zn^{2+} consists in adding sodium hydroxide to the solution containing the basic radical. All of these radicals are precipitated as the corresponding hydroxides**. Thus,

$$CaCl_2 + 2NaOH \rightarrow Ca(OH)_2\downarrow + 2NaCl$$
$$CuSO_4 + 2NaOH \rightarrow Cu(OH)_2\downarrow + Na_2SO_4$$
$$FeSO_4 + 2NaOH \rightarrow Fe(OH)_2\downarrow + Na_2SO_4$$
$$FeCl_3 + 3NaOH \rightarrow Fe(OH)_3\downarrow + 3NaCl$$
$$Pb(NO_3)_2 + 2NaOH \rightarrow Pb(OH)_2\downarrow + 2NaNO_3$$
$$ZnSO_4 + 2NaOH \rightarrow Zn(OH)_2\downarrow + Na_2SO_4$$

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*Any sulphate given to you for identification will be water-soluble.

**Calcium hydroxide is not precipitated if the solution is very fine.

The hydroxides of copper, iron (II), and iron (III) are readily recognized by their characteristic colours.

The hydroxides of calcium, lead, and zinc are white. However, the hydroxides of lead and zinc are amphoteric, and dissolve in the excess of sodium hydroxide solution.

$$\underset{\substack{\text{zinc}\\\text{hydroxide}}}{Zn(OH)_2} + \underset{\substack{\text{sodium}\\\text{hydroxide}}}{2NaOH} \rightarrow \underset{\substack{\text{sodium zincate}\\\text{(soluble)}}}{Na_2ZnO_2} + \underset{\text{water}}{2H_2O}$$

$$\underset{\substack{\text{lead}\\\text{hydroxide}}}{Pb(OH)_2} + \underset{\substack{\text{sodium}\\\text{hydroxide}}}{2NaOH} \rightarrow \underset{\substack{\text{sodium plumbate}\\\text{(soluble)}}}{Na_2PbO_2} + \underset{\text{water}}{2H_2O}$$

Therefore, if the white precipitate formed upon the addition of sodium hydroxide solution to the solution of the salt does not dissolve in excess alkali, it indicates calcium.

Method. The method of identifying basic radicals by adding sodium hydroxide solution involves the following steps :

(1) *Preparation of the solution of the salt.* Take about 0·2-0·3 g of the sample in a clean test tube. Add about 4-5 cm^2 of distilled water, and shake. If the salt does not dissolve, heat until the liquid boils. If the salt does not dissolve even in hot water, take a fresh lot of the salt in another test tube, and add 4-5 cm^3 of dilute hydrochloric acid. Shake thoroughly. If effervescence occurs, wait till it subsides. A clear solution should form. In some cases, solution may have to be prepared in dilute nitric acid.

Note tne colour of the solution obtained.

i. *Blue* colour of solution indicates *copper salt* or Cu^{2+} ions.

ii. *Light green* colour indicates *ferrous salt* or Fe^{2+} ions.

iii. *Yellow colour* indicates *ferric salt* or Fe^{3+} ions.

iv. Colourless solution indicates absence of these ions.

(2) *Addition of sodium hydroxide solution.* To the solution of the salt in a test tube, add sodium hydroxide solution from the reagent bottle, *dropwise.* Make observation with regard to the following :

(*i*) If a precipitate is formed or not.

(*ii*) If a precipitate is formed, what is its colour ?

(*iii*) Whether the precipitate dissolves in excess alkali or not.

(*iv*) Whether the colour of the precipitate is changed on heating or not.

(3) *Making deduction.* Make deductions based on observations according to Table 8-3.

Table 8-3. *Action of sodium hydroxide on the solution of the given salt*

	Observation	*Basic radical indicated*
1.	White precipitate consisting of very fine particles; insoluble in excess NaOH	Calcium, Ca^{2+}
2.	Light blue precipitate; insoluble in excess NaOH but turns black on boiling.	Copper, Cu^{2+}
3.	Light green, gelatinous precipitate; insoluble in excess NaOH.	Iron (II), Fe^{2+}
4.	Reddish-brown precipitate; insoluble in excess NaOH.	Iron (III), Fe^{3+}
5.	White precipitate that dissolves in the excess of sodium hydroxide solution.	Lead, Pb^{2+}, or Zinc, Zn^{2+}

5-1. Identification of NH_4^+. The addition of sodium hydroxide solution to the solution of an ammonium salt does not produce a precipitate. However, on warming ammonia is evolved. This gas can be readily recognized by its characteristic odour. Confirm its identity by the following tests :

(*i*) Hold moist red litmus paper over the mouth of the test tube; it goes blue.

(*ii*) Bring a glass rod moistened with concentrated hydrochloric acid near the mouth of the test tube; dense white fumes are produced.

6. Distinction Between Lead and Zinc

If on adding sodium hydroxide solution to the solution of the salt a white precipitate is formed that dissolves in the excess of sodium hydroxide solution, the salt may contain lead or zinc. To make distinction between the two, dissolve the given salt in water or dilute nitric acid, and divide the solution in two parts.

i. To one part of the solution add dilute hydrochloric acid. If a white precipitate is formed, the salt contains lead or Pb^{2+}. The precipitate is lead chloride. For example,

$$\underset{\text{lead nitrate}}{Pb(NO_3)_2} + 2HCl \rightarrow \underset{\substack{\text{lead chloride}\\\text{(white ppt.)}}}{PbCl_2\downarrow} + 2HNO_3$$

Add water, and heat. The precipitate of lead chloride will dissolve in hot water but reappear on cooling the solution.

The solution of a zinc salt will not form precipitate on addition of hydrochloric acid.

ii. To the other part of the solution add potassium iodide solution. If a yellow precipitate is formed, the salt contains lead or Pb^{2+}. For example,

$$\underset{\text{lead nitrate}}{Pb(NO_3)_2} + \underset{\text{potassium iodide}}{2KI} \longrightarrow \underset{\text{lead iodide}}{PbI_2\downarrow} + \underset{\text{potassium nitrate}}{2KNO_3}$$

No precipitate is formed in the case of zinc salts.

7. Additional Tests for Certain Cations

1. Calcium, Ca^{2+}. To the solution of the substance (in water, HCl, or HNO_3) add dilute sulphuric acid. A white precipitate of calcium sulphate, $CaSO_4$, is formed. For example,

$$CaCl_2 + H_2SO_4 \rightarrow CaSO_4\downarrow + 2HCl$$

2. Copper, Cu^{2+}. (1) To the solution of the substance add ammonium hydroxide, a few drops at a time and with shaking after each addition, till it is in excess. A light blue precipitate of copper hydroxide, $Cu(OH)_2$, is first formed which dissolves in excess NH_4OH yielding deep blue solution. For example,

$$\underset{\substack{\text{copper (II)}\\\text{sulphate}}}{CuSO_4} + \underset{\substack{\text{ammonium}\\\text{hydroxide}}}{2NH_4OH} \rightarrow \underset{\substack{\text{ammonium}\\\text{sulphate}}}{(NH_4)_2SO_4} + \underset{\substack{\text{copper (II) hydroxide}\\\text{(light blue ppt.)}}}{Cu(OH)_2\downarrow}$$

$$\underset{\text{light blue ppt}}{Cu(OH)_2} + \underset{\text{(in excess)}}{NH_4OH} \rightarrow \textit{Deep blue solution}$$

(2) Take 3-4 cm^3 of the solution of the substance in a test tube. Slip in a 2-3 cm long strip of magnesium ribbon into the solution. Allow to stand for a few minutes. The part of

the magnesium ribbon dipping in the solution will be covered with reddish-brown particles of copper displaced from the solution. For example,

$$\underset{\substack{\text{copper sulphate}\\ \text{(blue solution)}}}{CuSO_4} + \underset{\text{magnesium}}{Mg} \rightarrow \underset{\substack{\text{mag. sulphate}\\ \text{(colourless)}}}{MgSO_4} + \underset{\substack{\text{displaced copper}\\ \text{(covers magnesium)}}}{Cu\downarrow}$$

On shaking, the copper particles are detached and the reaction continues. Gradually the blue colour of the solution fades and when all copper has been displaced the solution becomes colourless.

Alternatively, clean iron filings may be used in place of magnesium. The iron also displaces copper from the solution which coats the filings.

$$CuSO_4 + Fe \rightarrow FeSO_4 + Cu\downarrow$$

The colour of the solution progressively fades. When all copper has been displaced, the solution becomes light green due to the dissolved $FeSO_4$.

3. **Ferrous iron, iron (II), Fe^{2+}.** To the solution of the substance add a few drops of concentrated nitric acid. Boil till brown fumes are no longer evolved. The original light green colour of the solution will change to yellow. This is because nitric acid oxidizes Fe^{2+} ions to Fe^{3+} ions which make yellow solution.

$$\underset{\text{(light green solution)}}{Fe^{2+}} \xrightarrow{\text{oxidize}} \underset{\text{(yellow solution)}}{Fe^{3+}}$$

To the yellow solution obtained add sodium hydroxide solution. A brown precipitate of $Fe(OH)_3$ will form and not the light green precipitate of $Fe(OH)_2$.

Alternatively, heat the solution with a few drops of dilute sulphuric acid and a little hydrogen peroxide. This will also cause the oxidation of Fe^{2+} to Fe^{3+}, and the solution will change colour from light green to yellow. On adding sodium hydroxide solution a brown precipitate of $Fe(OH)_3$ will form.

4. **Ferric iron, iron (III), Fe^{3+}.** To the solution of the substance add a few drops of sulphuric acid or hydrochloric acid* and a pinch of sodium sulphite or sodium bisulphite. Heat the mixture. The yellow colour of the original solution will change to light green. Cool, and add sodium hydroxide (or ammonium hydroxide). A light green precipitate of iron (II) hydroxide, $Fe(OH)_2$, will form and not the brown precipitate of $Fe(OH)_3$ as with the original solution.

Explanation. On heating with an acid and sodium sulphite (or bisulphite), the sulphur dioxide evolved reduces the Fe^{3+} ions in the solution to Fe^{2+} ions. As a result the colour of the solution changes from yellow to light green.

$$Na_2SO_3 + 2HCl \xrightarrow{\Delta} 2NaCl + H_2O + SO_2$$

$$\underset{\text{(yellow solution)}}{Fe^{3+}} \xrightarrow{\text{reduction by } SO_2} \underset{\text{(light green solution)}}{Fe^{2+}}$$

On adding sodium hydroxide solution, therefore, $Fe(OH)_2$ is precipitated and not $Fe(OH)_3$.

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*If the solution was prepared in hydrochloric acid, do not add any more acid; only add sodium sulphite or sodium bisulphite.

3. **Zinc, Zn^{2+}.** To the solution of the salt add ammonium hydroxide, a few drops at a time and with shaking, till it is in excess. A white precipitate of zinc hydroxide, $Zn(OH)_2$, is first formed which dissolves in the excess ammonium hydroxide forming a colourless solution. For example,

$$\underset{\text{zinc chloride}}{ZnCl_2} + \underset{\substack{\text{ammonium}\\\text{hydroxide}}}{2NH_4OH} \rightarrow \underset{\substack{\text{ammonium}\\\text{chloride}}}{2NH_4Cl} + \underset{\substack{\text{zinc hydroxide}\\\text{(white ppt.)}}}{Zn(OH)_2\downarrow}$$

$$Zn(OH)_2 + NH_4OH\ (\text{excess}) \rightarrow \textit{Colourless soluion}$$

Experiment 8-7. To find what happens when sodium hydroxide solution is added to solutions containing Cu^{2+}, Fe^{2+}, and Fe^{3+}.

Method. Add sodium hydroxide solution dropwise to each solution in turn. Add the alkali in excess after a precipitate has formed.

Observe the colour of the precipitate in each case and note that none of the three precipitates dissolves in excess alkali Record your observations, and write equations for each reaction.

Experiment 8-8. To find out which metal is present in the given substance A.

Method. Take a portion of the substance in a test tube, add water, and shake. The substance does not dissolve.

Take another portion of the substance in a test tube add dilute hydrochloric acid, and shake. The substance dissolves to give a clear solution.

To a part of the solution add sodium hydroxide solution dropwise. A white precipitate is formed which, however, dissolves in excess alkali. This shows that the metal in A is lead or zinc.

To another part of the solution of A add potassium iodide solution. No precipitate is formed. This shows that the metal in A is zinc.

Record your observations, and write equations for the reactions.

Experiment 8-9. To find the basic radical in the given substance C.

Method. Take a portion of the substance in a test tube, add water, and shake. The substance readily dissolves.

To the solution of the substance add sodium hydroxide solution. No precipitate is formed. This shows that the substance probably contains NH_4^+.

Heat the mixed solution. A colourless gas with a characteristic pungent odour is evolved.

Test the gas with moist red litmus paper which it turns blue and with a glass rod dipped in concentrated hydrochloric acid around which it produces dense white fumes.

These tests show that the gas evolved is ammonia.

Hence, the basic radical in the given substance A is NH_4^+.

Experiment 8-10. To identify the given salt A1.

Experiment	*Observation*	*Deduction*
1. Noted the physical characteristics of A1.	A1 is white and heavy.	A1 is probably a lead salt.
2. Treated A1 with dilute sulphuric acid.	No action in cold or on heating.	Carbonate and sulphide radicals absent.
3. Heated A1 with concentrated sulphuric acid.	A brown, heavy gas with a pungent smell is evolved.	**The acid radical in A1 is nitrate, NO^-_3.**
(*i*) Held a filter paper soaked in $FeSO_4$ solution near the mouth of the test tube.	The paper turns black	
(*ii*) Added copper turnings to the hot reaction mixture in the best tube.	Copious reddish-brown fumes are evolved.	
4. Applied flame test.	No characteristics coloration.	Na^+, K^+, and Ca^{2+} absent.
5. Dissolved A1 in water and added sodium hydroxide solution dropwise.	A white precipitate is formed which dissolves in excess alkali.	The basic radical in A1 is either Pb^{2+} or Zn^{2+}.
6. Added potassium iodide solution to the original solution of A1.	A yellow precipitate is formed.	The basic radical in A1 is Pb^{2+}.

Result. The given salt A1 is lead nitrate, $Pb(NO_3)_2$.

Experiment 8-10. To identify the given salt B1.

Experiment	*Observation*	*Deduction*
1. Noted the physical characteristics of B1.	B1 is a white crystalline substance.	Fe^{2+}, Fe^{3+}, and Cu^{2+} are absent.
2. Treated B1 with dilute sulphuric acid.	No action in cold or on heating.	Carbonate and sulphide radicals absent.
3. Heated B1 with concentrated sulphuric acid.	A colourless gas with pungent smell is evolved.	The gas evolved when B1 is heated with concentrated sulphuric acid is hydrogen chloride.
(*i*) Held moist blue litmus paper over the mouth of the test tube.	The litmus paper turns red.	The acid radical in B1 is chloride.
(*ii*) Brought a glass rod moistened with ammonia solution near the mouth of test tube.	Dense white fumes are produced.	
(*iii*) Dissolved a little of B1 in water and added silver nitrate solution.	A white, curdy precipitate is formed.	

4. Applied flame test.	No characteristic coloration.	Na^+, K^+, and Ca^{2+} absent.
5. Dissolved B1 in water and added sodium hydroxide solution.	No precipitate is formed.	Ca^{2+}, Pb^{2+}, Zn^{2+} are absent.
6. Heated the mixed solution.	A colourless gas with pungent smell is evolved.	The gas evolved when B1 is heated with sodium hydroxide solution is ammonia.
(*i*) Held a moist red litmus paper over the mouth of the test tube.	The litmus paper turns blue.	
(*ii*) Brought a glass rod dipped in ammonia solution near the mouth of the test tube.	Dense white fumes are produced.	∴ The basic radical in B1 is NH_4^+.

Result. The given salt B1 is ammonium chloride, NH_4Cl.

9

Acids, Bases, and Oxides

An acid is a substance which contains replaceable hydrogen, *i e.*, hydrogen which may be wholly or partly replaced by metals. According to the ionic theory, *an acid is a substance which yields hydrogen, ions, H^+, as the only positive charged ion in aqueous solution.* For example,

$$HCl \rightleftharpoons H^+ + Cl^-$$
$$H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}, \text{ etc.}$$

All the characteristic properties of acids are due to the H^+ ions they yield in solutions.

A base is the hydroxide of a metal and reacts with acids to form salts and water. Sodium hydroxide, $NaOH$; potassium hydroxide, KOH; zinc hydroxide, $Zn(OH)_2$, etc., are all examples of bases. The soluble bases, *e.g.*, sodium hydroxide and potassium hydroxide are called *alkalis*. *The alkalis yield a high concentration of hydroxyl ions, OH^-, in aqueous solutions.* For example,

$$NaOH \rightleftharpoons Na^+ + OH^-$$
$$KOH \rightleftharpoons K^+ + OH^-$$

The characteristic properties of alkalis are due to the OH^- ions they yield in solutions.

Oxides are the binary compounds oxygen. The oxides of metals may be divided into two important classes : (1) the normal oxides, *e.g.*, copper oxide, CuO, and (2) the dioxides, *e.g.*, manganese dioxide, MnO_2.

1. Characteristic Properties of Acids

1. Acids turn litmus red, methyl orange pink, and decolorize phenolphthalein solution made pink with a drop of alkali.

2. Acids cause effervescence with sodium carbonate (or, any other carbonate) giving off a colourless gas which is carbon dioxide. This gas may be tested by passing it through lime water which is turned milky.

3. Acids react with magnesium, giving off a colourless gas which is hydrogen. This gas may be tested by bringing a lighted splint near it; the splint is extinguished and the gas burns with a 'pop'.

These properties may be used to determine whether a given solution is an acid or not.

2. Characteristic Properties of Alkalis

1. Alkalis turn litmus blue, methyl orange solution yellow, and phenolphthalein solution pink.
2. They neutralize acids to form salts and water.
3. They do not react with sodium carbonate and magnesium.

The solution of an alkali can be distinguished from the solution of an acid on the basis of these characteristics.

Experiment 9-1. To determine which one of the two given solutions X and Y is acidic and which is basic.

Experiment	*Observations*	*Deductions*
1. Tested each one of the two solutions with red and blue litmus papers.	X turns blue litmus red. Y turns red litmus blue.	X is an acid. Y is an alkali.
2. Tested each solution with phenolphthalein solution.	No change with X. Pink colour with Y.	X is an acid. Y is an alkali.
3. Added a pinch of sodium carbonate to a test portion of each solution in a test tube.	Brisk effervescence occurs with X. No action with Y.	X is an acid. Y is an alkali.
4. Added magnesium powder to a test portion of each solution.	Brisk effervescence with X (due to the liberation of hydrogen). No action with Y.	X is an acid. Y is an alkali.

Results. Solution X is an acid ; solution Y is an alkali.

3. Oxides

The normal oxides of metals, *e.g.*, copper oxide (CuO), zinc oxide (ZnO), etc., dissolve in acids to form corresponding salts and water. For example,

$$CuO + 2HCl \rightarrow CuCl_2 + H_2O$$
$$ZnO + H_2SO_4 \rightarrow ZnSO_4 + H_2O$$

The dioxides, such as manganese dioxide, MnO_2, and lead dioxide, PbO_2, too, dissolve in acids to form corresponding salts and water. However, in addition they liberate chlorine from hot, concentrated hydrochloric acid and oxygen from hot, concentrated sulphuric acid.

$$MnO_2 + 4HCl \xrightarrow{\Delta} MnCl_2 + 2H_2O + Cl_2\uparrow$$
$$PbO_2 + 4HCl \xrightarrow{\Delta} PbCl_2 + 2H_2O + Cl_2\uparrow$$
$$2MnO_2 + 2H_2SO_4 \xrightarrow{\Delta} 2MnSO_4 + 2H_2O + O_2\uparrow$$

Thus, distinction between an oxide and a dioxide can easily be made by heating the sample with concentrated hydrochloric acid. If the sample dissolves to form the solution of a salt without the evolution of chlorine, it is a normal oxide; if chlorine is evolved, it is a dioxide. Chlorine is recognized by the following tests :

(*i*) Greenish-yellow colour and pungent, choking smell.

(*ii*) It turns blue litmus first red and then colourless.

(*iii*) It produces deep blue coloration on starch-iodide paper.

[Starch-iodide paper is made by dipping a strip of filter paper in a mixed solution of potassium iodide and starch. Chlorine liberates iodine from potassium iodide and the liberated iodine produces deep blue colouration with starch.

$$2KI + Cl_2 \rightarrow 2KCl + I_2]$$

Experiment 9-2. To determine which one of the two given samples A and B is copper oxide and which one manganese dioxide.

Method. Take a portion of sample A in a clean, dry test tube in a holder. Add concentrated hydrochloric acid to the sample, and heat. Observe what happens.

Repeat the test with a portion of sample B.

Observations. Sample A dissolves forming a blue solution without the evolution of a gas.

Sample B liberates a greenish-yellow gas with a pungent, choking smell. The gas turns moist blue litmus paper first red and then colourless, and produces deep blue coloration in starch-iodide paper. Hence, this gas is chlorine.

Deductions. Sample A is copper oxide; the blue solution formed on heating it with concentrated hydrochloric acid contains copper (II) chloride, $CuCl_2$.

$$\underset{\text{copper (II) oxide}}{CuO} + \underset{\text{hydrochloric acid}}{2HCl} \xrightarrow{\Delta} \underset{\text{copper (II) chloride}}{CuCl_2} + \underset{\text{water}}{H_2O}$$

Sample B is manganese dioxide; the gas evolved when it is heated with concentrated hydrochloric acid is chlorine.

$$\underset{\text{manganese dioxide}}{MnO_2} + \underset{\text{hydrochloric acid}}{4HCl} \xrightarrow{\Delta} \underset{\text{manganese chloride}}{MnCl_2} + \underset{\text{water}}{2H_2O} + \underset{\text{chlorine}}{Cl_2\uparrow}$$

10

Percentage Composition of Mixtures

A mixture is a simple collection of two or more substances each one of which keeps its own properties. The constituents of a mixture can, therefore, be separated by physical methods based on such properties as solubility in various solvents, volatility, etc. The percentage composition of a given mixture can be determined by taking weighed amount of the mixture, separating is constituents, and weighing each constituent.

In this chapter we shall study the method for the determination of the percentage composition of mixtures of two constituents one of which is soluble in water and the other one is insoluble.

1. Percentage Composition of a Mixture of Powdered Salt and Sand

The percentage composition of a mixture of powdered salt and sand can be determined in the following steps :

1. Weighing out a certain quantity of the mixture.
2. Dissolving out the salt by stirring the weighed mixture with water.
3. Separating the undissolved sand by filtration.
4. Drying the sand thus separated.
5. Weighing the sand.

Experiment 10-1. To determine the percentage composition of the given mixture of powdered common salt and washed sand

Requirement. Beakers, funnel stand, filter paper, watch glass, pair of tongs, glass rod, drying cone, sand bath, chemical balance, weight box, etc.

Theory. Of the two components of the given mixture, sodium chloride is soluble in water while sand is insoluble. Hence, sodium chloride can be removed from the mixture by dissolving out in water. The undissolved sand can be obtained by filtration as the insoluble residue. It is dried and weighed. From the weight of the mixture taken and the weight of *dry* sand separated from it, the percentage composition of the mixture can be calculated.

Method. Proceed in the following steps :

1. *Weigh out about 5-6 g of the given mixture*. Take a clean, dry watch glass and weigh it. Remove the watch glass from the balance pan, put about 5-6 g of the mixture in it, and weigh again.

The difference in the two weights is the weight of the mixture taken.

2. *Transferring the mixture to a beaker.* Take a clean beaker and carefully transfer into it the mixture from the watch glass. Wash down any particles adhering to the watch glass into the beaker by directing a jet of distilled water from the wash bottle in the watch glass.

3. *Dissolving out sodium chloride.* Add a further quantity of distilled water to the mixture in the beaker. Stir thoroughly with a glass rod. Warm the contents of the beaker for a while and than allow to stand. Almost whole of the sodium chloride in the mixture will dissolve in water and the undissolved sand will settle down to the bottom.

4. *Separating sand by filtration.* Take a filter paper, and place it in the left pan of the balance. Counterpoise it with another filter paper (making suitable adjustment, if necessary), and leave it there. Remove the filter paper from the left pan, fold it into a cone and fit it in a funnel. Moisten the filter paper cone with water and place the funnel in a conical flask. Decant the supernatant liquid from the beaker into the filter paper cone with the help of a glass rod, taking care that the layer of sand at the bottom is not disturbed. Pour another lot of 20-30 cm^3 of distilled water in the beaker, stir the contents with a glass rod, allow to stand until the sand settles down, and decant the clear supernatant liquid as before. Repeat '*washing by decantation*' until a test portion of the filtrate no longer produces turbidity with silver nitrate solution. This test shows that sodium chloride (which gives a white precipitate of silver chloride with silver nitrate solution) has been completely removed and the residue of sand is absolutely free from it.

Now transfer sand to the filter along a glass rod by playing a jet of water from the wash bottle. Make sure that whole of the sand has been washed down to the filter paper and

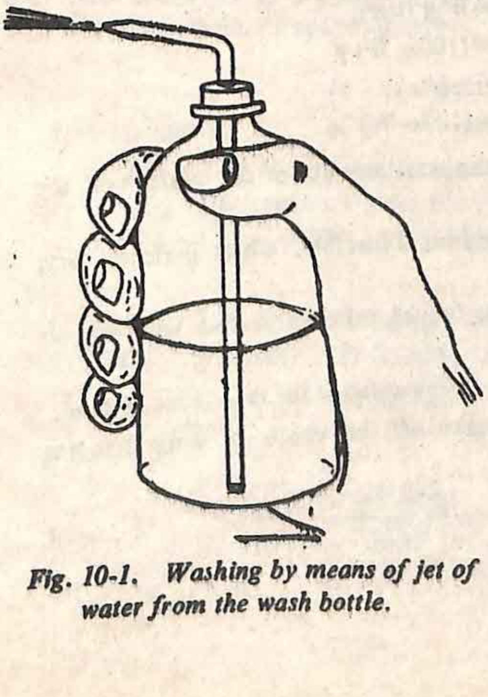

Fig. 10-1. Washing by means of jet of water from the wash bottle.

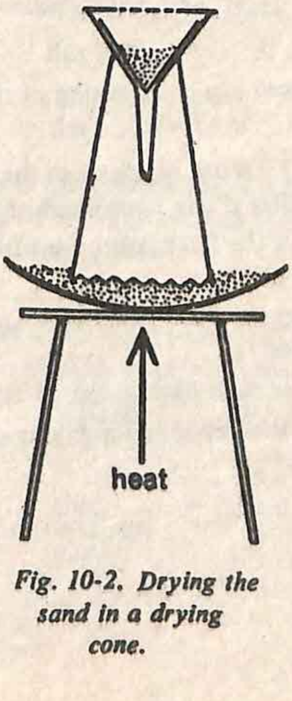

Fig. 10-2. Drying the sand in a drying cone.

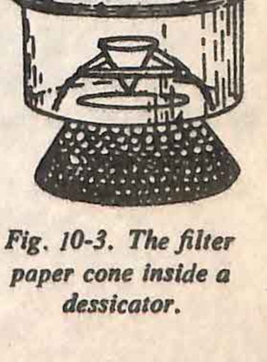

Fig. 10-3. The filter paper cone inside a dessicator.

not a single particle has been left behind. After the water in the filter paper cone has drained out, wash down the sand particles sticking to the upper edge into the central cavity.

5. *Drying the sand in the filter-paper cone.* Cover the funnel with a piece of paper, press its extra portion round the edge of the funnel, and make a few holes in the centre with a pointed glass rod. Place the covered funnel in a drying cone standing in a sand bath supported over a tripod stand (Fig. 10-2). Continue heating until the filter paper and the sand are perfectly dry. Now, carefully remove the filter-paper cone from the funnel, support it on a clay pipe triangle. and place in a dessicator for cooling (Fig. 10-3).

6. *Washing the sand.* When cool, place the filter-paper cone containing the sand resting slantingly on the edge of the left pan of the balance with a counterpoising filter paper in the right pan. Place sufficient weights on the right pan to counterbalance. These weights represent the weight of the sand in filter paper cone :

Observations

1.	Weight of the empty watch glass	$=x$ g
2.	Weight of the watch glass and mixture	$=y$ g
3.	Weight of the sand	$=z$ g

Calculations

Weight of the mixture taken for the experiment $=(y-x)g$

Weight of the sand in the mixture $=z$ g

$\therefore$ $(y-x)g$ of the mixture contains sand $=z$ g

$\therefore$ 100 g of the mixture contains sand $=\dfrac{z\text{ g}}{(y-x)\text{ g}}\times 100$ g

$=W$ g (say)

and salt $=(100-W)$ g

Result. The percentage composition of the mixture is :

sand $=$ W% ; salt $=(100-\text{W})$ %.

Precautions. (1) Work carefully so that not the least amount of the mixture or the sand is lost at any stage after it has been weighed out.

(2) Do not pull out the filter-paper cone from the funnel forcibly; when perfectly dry, it can be slipped out easily.

(3) While removing the filter-paper cone from the funnel, take care that your fingers do not touch sand particles.

(4) Start heating the sand bath about 15 minutes before using it for drying moist sand.

(5) Before placing the funnel in the drying cone shake off the drops of water hanging by the stem of the funnel.

APPENDIX I

Solved Experimental Exercises

Exercise 1

(a) Take a small portion of the substance A1 in a clean test tube and add 2 cm^3 *of dilute sulphuric acid. Warm gently till a clear solution is obtained. Filter and divide the solution in two parts.*

i. Write down the observations made when A1 is heated with dilute sulphuric acid. Make any deduction you can.

ii. To one part of the filtrate add sodium hydroxide solution in excess. Warm gently. Write down your observations and the deduction from the observations made.

iii. In the other part of the filtrate place about 2 cm long strip of magnesium. Write down your observation and deduction. Write chemical equations for the reactions involved.

i. **Observations**

1. When A1 is heated with dilute sulphuric acid, it dissolves in it forming a blue solution, but no gas is evolved.

Deduction

The blue colour of the solution shows that A1 is a compound of copper. Since no gas is evolved, it is oxide of copper, CuO, which is confirmed by the black colour of A1. The blue solution contains copper sulphate.

$$CuO + H_2SO_4 \xrightarrow{\Delta} CuSO_4 \quad H_2O$$

ii. **Observations**

1. On adding sodium hydroxide solution a blue precipitate is formed.
2. The blue precipitate does not dissolve in excess of sodium hydroxide.
3. On warming gently, the blue precipitate turns black.

Deduction

The filtrate contains Cu^{2+} ions which form a blue precipitate of copper hydroxide on adding sodium hydroxide. On heating copper hydroxide changes to black copper (II) oxide.

$$\underset{\text{blue filtrate}}{CuSO_4} + 2NaOH \longrightarrow \underset{\text{blue precipitate}}{Cu(OH)_2\downarrow} + Na_2SO_4$$

$$\underset{\substack{\text{copper hydroxide}\\\text{(blue)}}}{Cu(OH)_2} \xrightarrow{\Delta} \underset{\substack{\text{copper (II) oxide}\\\text{(black)}}}{CuO} + H_2O$$

iii. **Observations**

1. Magnesium strip is slowly coated with red particles of copper.
2. The blue colour of the filtrate slowly fades and finally becomes colourless.

Deduction

The blue filtrate contains Cu^{2+} ions which are reduced to copper atoms by magnesium. With the displacement of copper from the solution, its blue colour fades.

$$\underset{\text{blue solution}}{CuSO_4} + Mg \longrightarrow \underset{\text{colourless solution}}{MgSO_4} + \underset{\text{red particles}}{Cu\downarrow}$$

(b) i. Take a pinch of A2 in a clean, dry, hard-glass test tube. Heat strongly and record your observations. If a gas is evolved, identify it and make deduction about the identity of A2.

ii. Add a pinch of A2 to 2 cm³ of concentrated hydrochloric acid taken in a test tube. Warm gently. Identify the gas evolved with suitable tests. Write down your observations and make deduction regarding the chemical nature of A2.

i. **Observations.**

(1) The chocolate-brown A2 changes to a yellow-orange powder.

(2) A colourless gas is evolved which relights a glowing splint.

Deduction

A2 is lead dioxide. It gives off oxygen when heated.

ii. **Observations**

(1) A greenish-yellow gas is evolved on warming A2 with concentrated hydrochloric acid.

(2) The greenish-yellow gas turns moist blue litmus paper first red and then colourless.

(3) The gas produces deep blue coloration in starch-iodide paper.

From these tests the gas is identified to be chlorine.

Deduction

A2, *i.e.*, lead dioxide, is an oxidizing agent, and it oxidizes hydrochloric acid to chlorine.

$$PbO_2 + 4HCl \xrightarrow{\Delta} PbCl_2 + 2H_2O + Cl_2\uparrow$$

Exercise 2

(a) Take a portion of substance B1 in a dry test tube and heat it. Record your observations and identify the gas evolved.

Observations

1. When B1 is heated, it decrepitates.
2. A heavy, brown gas is evolved.

Gas identified. Nitrogen dioxide, NO_2

(b) Take another portion of B1 in a dry test tube. Add about 3-4 cm³ of distilled water and warm. Filter, if necessary, to get a clear solution. Divide the solution in two parts.

i. To one part of the solution add dilute hydrochloric acid in cold. Then add 5 cm^3 of distilled water, and heat again. Note down the observations and make your deduction about the substance formed in this experiment.

ii. To the other part of the solution add a few drops of potassium chromate solution. Write down your observation and deduction.

iii. Deduce the identity of substance B1, and write equations for the reactions in (a) and (b).

i. Observations

1. A curdy, white precipitate is formed on adding hydrochloric acid.
2. The precipitate dissolves when heated with water, but appears again on cooling.

Deduction

As the white precipitate dissolves on heating with water, only to reappear on cooling, it must be lead chloride, $PbCl_2$.

ii. Observation

A thick, yellow precipitate is formed on adding potassium chromate solution.

Deduction

The solution contains Pb^{2+} ions which form a yellow precipitate of lead chromate, $PbCrO_4$, with potassium chromate.

iii. Substance B1 is a nitrate as it gives off nitrogen dioxide when heated. Reactions under (*b*) show that the basic radical in B1 is lead. Hence, B1 is lead nitrate.

Equations

1. $$\underset{\text{lead nitrate}}{2Pb(NO_3)_2} \xrightarrow{\Delta} \underset{\text{lead oxide}}{2PbO} + \underset{\text{nitrogen dioxide}}{4NO_2} + \underset{\text{oxygen}}{O_2\uparrow} \qquad \text{[Action of heat on B1]}$$

2. $$\underset{\text{lead nitrate}}{Pb(NO_3)_2} + 2HCl \rightarrow \underset{\substack{\text{lead chloride}\\ \text{(white ppt.)}}}{PbCl_2\downarrow} + \underset{\text{nitric acid}}{2HNO_3} \qquad \begin{matrix}\text{[Formation of white ppt.}\\ \text{on adding hydrochloric acid]}\end{matrix}$$

3. $$\underset{\text{lead nitrate}}{Pb(NO_3)_2} + K_2CrO_4 \rightarrow \underset{\substack{\text{lead chromate}\\ \text{(yellow ppt.)}}}{PbCrO_4\downarrow} + \underset{\substack{\text{potassium}\\ \text{nitrate}}}{2KNO_3} \qquad \begin{matrix}\text{[Formation of yellow ppt.}\\ \text{on adding ppt. chromate]}\end{matrix}$$

Exercise 3

(*a*) *Take a pinch of the substance A in a clean. dry test tube and heat strongly. Apply suitable tests to identify the gas evolved, and record all observations.*

Cool the residue and keep it for further tests (b) *and* (b) *that will follow.*

Observations

1. A colourless gas is evolved.
2. The gas has no smell.
3. A lighted splint introduced in the gas is extinguished.

4. A drop of clear lime water held on a glass rod near the mouth of the test tube turns milky.

(*b*) *Transfer a part of the residue obtained in* (*a*) *to another test tube. Dissolve it in the minimum quantity of dilute hydrochloric acid. To the solution obtained add sodium hydroxide solution in small lots till the alkali is in excess (test with red litmus paper). Note down your observations Identify the cation present in A and give reasons for the deduction made by you.*

Observations

1. On adding sodium hydroxide solution, a white precipitate consisting of very fine particles is formed.

2. The precipitate does not dissolve in the excess of sodium hydroxide solution.

Deduction

The cation present in A is Ca^{2+}. The reasons for this deduction are that (1) the precipitate consists of very fine particles and (2) it does not dissolve in the excess of the alkali.

(*c*) *To the solution obtained in* (*b*) *add dilute sulphuric acid till there is no further change. Note down your observation.*

Observation

On adding sulphuric acid to the solution obtained in (*b*), a white precipitate is formed.

(*d*) *To the remaining residue obtained in* (*a*) *add a concentrated solution of ammonium sulphate, and heat. Identify the gas evolved on the basis of your observations.*

Observations

1. The gas evolved is colourless.

2. It has a characteristic pungent odour.

3. It turns moist red litmus paper blue.

4. It produces dense white fumes in contact with a glass rod moistened with concentrated acid.

On the basis of these observations, the gas is identified as ammonia.

(*e*) *Name the compound A on the basis of the above experiments.*

Name of the Compound

The name of the compound A is calcium carbonate, $CaCO_3$.

Exercise 4

(*a*) *Take a portion of the substance B7 in a clean, dry, hard-glass test tube. Heat gently at first and then strongly until no further change occurs Identify the gas evolved on the basis of your observations, and also note down the colour of the residue. Make deduction with regard to the negative ion in the compound.*

Keep the residue for experiments (*b*) *and* (*c*).

Observations

1. When B7 is heated, a heavy, dark brown gas is evolved.
2. The gas has a pungent smell.
3. The gas turns a light green solution of iron (II) sulphate black.
4. The colour of the residue in the test tube is black.

Deduction

The compound B7 contains nitrate, NO_3^-, as the negative ion.

(*b*) *Take about one-half of the residue in a test tube. Add a little concentrated hydrochloric acid and heat. Cool, add water, and filter. To the filtrate add ammonium hydroxide solution in small lots till it is in excess. Record your observations.*

Observations

1. On adding ammonium hydroxide, a light blue precipitate is formed.
2. The precipitate dissolves in the excess of ammonium hydroxide solution yielding a deep blue solution.

(*c*) *Transfer the remaining residue in another test tube. Add a small amount of dilute sulphuric acid and heat till a clear solution is formed. To the solution thus obtained add sodium hydroxide solution till the alkali is in excess. Heat the contents of the test tube. Record the observations and identify the metallic ion present. Give two reasons for your deduction.*

Observations

(1) On adding sodium hydroxide solution, a light blue precipitate is formed.
(2) The light blue precipitate does not dissolve in the excess of alkali.
(3) On heating the contents, the light blue precipitate turns black.

Identification of metallic ion

The solution contains the metallic ion Cu^{2+}.

Reasons for deduction

(1) Copper compounds are blue, and the blue precipitate formed on addition of sodium hydroxide must be copper (II) hydroxide, $Cu(OH)_2$.

(2) On heating, copper hydroxide changes to copper (II) oxide which is black in colour.

$$\underset{\text{blue}}{Cu(OH)_2} \xrightarrow{\Delta} \underset{\text{black}}{CuO} + H_2O$$

Exercise 5

(*a*) *Take about one half of the substance B8 in a dry test tube. Heat strongly for 5-6 minutes. Write down the observations and identify the gas given off. Make deductions about the identity of the gas and the negative ion in B8.*

Observations

1. A colourless gas is evolved.
2. The gas has no smell.

3. A lighted splint introduced into the gas is extinguished
4. The gas turns lime water milky.
5. The residue is yelow when hot, and white when cold.

Deductions

1. The gas evolved is carbon dioxide.
2. The negative ion in B8 is carbonate, CO_3^{2-}.

(*b*) *Take the remaining quantity of B8 in another test tube. Add dilute sulphuric acid. Record all your observations and identify the gas evolved. Keep the solution for experiments* (*c*) *and* (*d*).

Observations

1. There is a sudden effervescence and bubbles of a colourless add odourless gas are formed.
2. The gas extinguishes a lighted splint and turns lime water milky.

Identification of the gas

The gas evolved is carbon dioxide, CO_2.

(*c*) *Divide the solution obtained in* (*b*) *in two parts. To one part add sodium hydroxide solution in small lots till the alkali is in excess. Record the changes observed. What inference can you draw from the experiment ? Give reasons.*

Observations

1. First, a white gelatinous precipitate is formed.
2. The precipitate dissolves in the excess of alkali to yield a clear solution.

Inference

The solution contains Zn^{2+} ions. These ions produce a white precipitate of $Zn(OH)_2$ on addition of alkali. This precipitate is soluble in the excess of alkali. This inference is supported by the fact that the residue in (*a*) is yellow when hot and white when cold.

(*d*) *To the other part of the solution obtained from* (*b*) *slowly add ammonium hydroxide till is excess. Write down your observations.*

Observations

1. A white precipitate is formed on addition of ammonium hydroxide.
2. The precipitate dissolves in the excess of ammonium hydroxide to yield a clear solution.

(*e*) *From the experiments conducted name the compound and write its formula.*

Name of the compound. Zinc carbonate.

Formula of the compound. $ZnCO_3$.

Exercise 6

Divide the substance B16 in nearly four equal parts.

(*a*) *Take one part of the substance B16 in a dry, hard-glass test tube. Heat, first gently and then strongly for about five minutes. Record your observations and state the nature of the gases evolved on strong heating. Identify the residue left. Make deduction about the class of compounds to which B16 belongs.*

Observations

1. On gentle heating, a colourless and odourless vapour is evolved. The vapour condenses on the upper part of the test tube as drops of a colourless liquid. This liquid turns anhydrous copper sulphate blue.
2. On strong heating, a colourless gas with a pungent smell is evolved.
3. The pungent smelling gas turns moist blue litmus paper blue. So, it is acidic in nature.
4. The residue left behind is black when hot and rust-red when cold.

Identification of residue.

The residue is iron (III) oxide, Fe_2O_3.

Deduction.

The substance B16 belongs to the class of hydrated compounds.

(*b*) *Dissolve another part of the substance B16 in distilled water. To the solution add sodium hydroxide solution in small lots till it is in excess. Write down your observations and the inference drawn.*

Observations

1. The solution of B16 in water is light green in colour.
2. When sodium hydroxide is added to the solution, a light green precipitate is formed.
3. The precipitate does not dissolve in the excess of sodium hydroxide.

Inference

The substance B16 is a compound of iron and contains Fe^{2+} ions. These ions produce a light green precipitate of $Fe(OH)_2$ on addition of sodium hydroxide.

(*c*) *Dissolve the third part of the substance B16 in about 4 cm^3 of dilute sulphuric acid. Add about 4cm^3 of hydrogen peroxide, heat gently, and then cool. Now add ammonium hydroxide till it is in excess. Note down your observations and inference.*

Observations

1. On warming with hydrogen peroxide, the solution of B16 in dilute sulphuric acid changes its colour from light green to yellow.
2. On adding ammonium hydroxide, a brown precipitate is formed, which does not dissolve in the excess of the alkali.

Inference

The Fe^{2+} ions in the solution of B16 are oxidized by hydrogen peroxide to Fe^{3+} ions. The Fe^{3+} ions produce a brown precipitate of $Fe(OH)_3$ on adding ammonium hydroxide.

(*d*) *Dissolve the remaining substance B16 in distilled water. To the solution add barium chloride and then dilute hydrochloric acid. Write down the observations and inference.*

Observations

1. On adding barium chloride solution, a thick, white precipitate is formed.
2. The precipitate does not dissolve in dilute hydrochloric acid.

Inference

The substance B16 contains sulphate, SO_4^{2-}, as the acid radical. This radical gives a white precipilate of barium sulphate upon addition of barium chloride.

$$BaCl_2+SO_4^{2-}\rightarrow BaSO_4\downarrow+2Cl^-$$

(*e*) *Write down the name and the formula of the substance of B16*

Name of B16. Iron (II) sulphate heptahydrate.

Formula of B16 $FeSO_4.7H_2O$

Exercise 7

(*a*) *Take about 2-3 cm³ of dilute* ***sulphuric acid*** *in a clean test tube, and add to it 2 pieces of* B_1. *Write down your observations and identify the gas evovled. Draw inference with regard to the nature of B1 based on reasoning.*

Observations

1. A sudden effervescence takes place and tiny bubbles of a colourless gas are formed.
2. The gas evolved has no smell.
3. When a lighted splint is brought near the mouth of the test tube, it is extinguished but the gas burns with a pop.

   These observations show that the gas evolved is hydrogen.
4. B1 slowly dissolves in the acid forming a colourless solution.

Deduction

Since B1 liberates hydrogen from dilute sulphuric acid, it must be an active metal. However, it could not be iron as it forms a colourless solution in acid.

(*b*) *Filter the contents obtained in* (*a*) *and divide the filtrate in two parts in clean test tubes. To one part of the filtrate add sodium hydroxide solution, a few drops at a time with shaking, until it is in excess. To the other part add sodium carbonate till a change occurs. Write down your observations and deduction.*

Observations

1. On adding sodium hydroxide to the solution of B1 in dilute sulphuric acid, a white precipitate is formed which dissolves in the excess of the alkali.
2. On adding sodium carbonate, a white precipitate is formed.

Deduction

The solution of B1 in dilute sulphuric acid contains Zn^{2+} ions. These ions produce a white precipitate of $Zn(OH)_2$ with sodium hydroxide ; $Zn(OH)_2$ dissolves in the excss of the alkali. On adding sodium carbonate, Zn^{2+} ions form a white precipitate of $ZnCO_3$.

(*c*) *From the above experiments, identify the substance B1.*

Identification of B1

B1 is zinc metal

Experiment 8

(*a*) *Place a small amount of the substance B2 in a dry, clean, hard-glass test tube. Heat it, first gently and then strongly for about 5 minutes. Record all your observations and make appropriate inference.*

Keep the residue for experiments (*b*), (*c*), (*d*), *and* (*e*).

Observations

1. A colourless and odourless vapour is evolved.
2. The vapour condenses as drops of a colourless and odourless liquid on the upper part of the test tube.
3. The condensed liquid turns anhydrous copper sulphate blue, showing that it is water.

Inference

The substance B2 contains water of hydration which is given off when it is heated.

(*b*) *Take a part of the residue from* (*a*) *in a test tube and add dilute hydrochloric acid. Record the observations and identify the gas evolved. What inference can you draw from the experiment ?*

Observations

1. There is a sudden effervescence and bubbles of a colourless gas are formed.
2. The colourless gas evolved has no smell.
3. A lighted splint brought near the mouth of the test tube is extinguished.
4. A drop of clear lime water at the end of a glass rod held at the mouth of the test tube becomes milky.

The observations show that the gas evolved is carbon dioxide, CO_2.

Inference

The substance B2 is a carbonate as it gives off carbon dioxide when treated with dilute hydrochloric acid.

(*c*) *Dissolve a part of the residue* (*or a little of the original substance*) *in distilled water. To the solution add a drop of methyl orange. Note down your observation and inference.*

Observation

The solution acquires a yellow colour.

Inference

The substance B2 is alkaline in nature.

(c) *Mix a part of the residue from (a) or the original substance with about an equal quantity of ammonium sulphate. Moisten the mixture with water, and heat. Identify the gas produced.*

Identification of the gas

1. The gas evolved is colourless, with a characteristic pungent smell.
2. A moist red litmus paper brought near the gas is turned blue.
3. Dense, white fumes are produced when a glass red moistened with concentrated hydrochloric acid is brought near the gas. From these observations, the gas is identified as ammonia, NH_3.

(e) *Carry out Flame test with B2 or its residue from experiment (a). Describe the test and make a deduction with regard to the metallic radical or ion present in B2.*

Description of flame test

Took a platinum wire sealed at one end in a glass tube. Dipped the free end of the wire in concentrated hydrochloric acid placed on a watch glass, and brought it into the colourless Bunsen flame. Repeated the dipping of the wire in the acid and placing it the flame till it ceased to impart any colouration to the flame.

Placed a fresh lot of concentrated hydrochloric acid on a watch glass. Dipped the free end of the wire in the acid and touched with it the substance B2. A tiny particle of B2 stuck to the moistened wire. Brought the wire into the non-luminous Bunsen flame which acquired a bright golden-yellow colour.

Deduction

B2 contains the metallic radical sodium, Na^+.

(*f*) *Write down the name and the compound B2.*

Name of B2

The compound B2 is hydrated sodium carbonate.

Exercise 9

(*a*) *Take a pinch of the substance B7 in a clean test tube. Add a small quantity of dilute hydrochloric acid, and warm. Write down your observations in identifying the gas evolved. What inference can you draw from the experiment?*

Observations

1. A colourless gas with the smell of burning sulphur is evolved.
2. The gas turns red a moist blue litmus paper held near the mouth of the test tube.
3. A filter paper strip dipped in potassium dichromate solution becomes green when exposed to the gas.

Deduction

The substance B7 gives off sulphur dioxide gas when warmed with dilute sulphuric acid. Hence, it is a sulphite.

(*b*) *Perform flame test with B7. Describe the test. Write down your observation and the deduction on the basis of the same.*

Description of flame test

As in Exercise 9.

Observation

The substance B7 imparts a golden yellow coloration to the flame.

Deduction

The substance B7 contains the metallic radical sodium, Na^+.

(*c*) *Write down the name and the formula of B7.*

Name of B7

The name of the substance B7 is sodium sulphite.

Formula of B 7

The formula of the substance B7 is Na_2SO_3.

Exercise 10

The given substance A20 is a solution containing iron (III) chloride and concentrated hydrochloric acid supplied to you in a test tube.

(*a*) *Take 2 cm³ of the substance A20 in a test tube. Add sodium hydroxide solution slowly till the alkali is in excess. Write down your observations and make your deduction on their basis.*

Observations

1. On adding hydrochloric acid, a brown precipitate appears which dissolves on shaking.
2. A permanent brown precipitate is formed when the alkali has been added in excess.

Deduction

The solution contains acid and Fe^{3+} ions. The acid neutralizes the alkali in the beginning. When the alkali is in excess, brown precipitate of $Fe(OH)_3$ is formed.

(*b*) *Take about 2 cm³ of the substance A20 in a test tube. Add a few drops of silver nitrate solution followed by concentrated nitric acid. Record your observations and the inference drawn.*

Observations

1. On adding silver nitrate solution a curdy, white precipitate is formed.
2. The precipitate does not dissolve in concentrated nitric acid.

Inference

The solution contains chloride ions, Cl^-.

(*c*) *Take 2-3 cm³ of the substance A 20 in a test tube and add a small piece of granulated zinc. Note down the observations and identify the gas evolved.*

Observations

1. There is a sudden effervescence on adding zinc and tiny bubbles of a colourless gas are formed. The gas has no odour.

2. When a lighted splint is brought near the gas, it burns with a pop; the splint is extinguished.

Gas identified

The gas evolved in the reaction is hydrogen.

(*d*) *Take 4 cm^3 of the substance A20 in a test tube. Add to it about 0.5 g of sodium sulphite and heat gently for 2-3 minutes. Cool the solution and add to it sodium hydroxide solution till it is in excess. Record your observations and the inference drawn.*

Observations

1. When A20 is heated with sodium sulphite, a colourless gas with the smell of burning sulphur is evolved.
2. The solution changes its yellow colour to light green.
3. On adding sodium hydroxide to the solution a light green precipitate is formed.

Inference

1. A20 contains acid which reacts with sodium sulphite to liberate sulphur dioxide gas.
2. Sulphur dioxide gas reduces Fe^{3+} ions in the solution to Fe^{2+} ions causing the solution to become light green.
3. On adding sodium hydroxide, a light green precipitate of iron (II) hydroxide is formed.

Exercise 11

(*a*) *Mix a pinch of the substance B10 with about an equal quantity of powdered slaked lime. Place the mixture in a clean hard-glass test tube and heat. Record your observations and describe the identification of the gas. Make a deduction about the cation present in B 10 and a give reason for it.*

Observations

1. When the mixture is heated, a colourless gas is evolved.
2. The gas has a characteristic, pungent odour.

Identification of gas

1. Moist red litmus paper was held near the mouth of the test tube ; it turned blue.
2. A glass red moistered with concentrated hydrochloric acid was brought near the mouth of the test tube. Dense, white fumes were produced.

The gas evolved is, therefore, ammonia, NH_3.

Deduction

Since the substance B10 gives off ammonia when heated with slaked lime, the cation present in it is NH_4^+ (or, B10 is an ammonium salt).

(*b*) *Dissolve some of the substance B10 in distilled water, and divide the solution in two parts :*

(*i*) *To one part of the solution add barium chloride solution followed by concentrated hydrochloric acid. Write down your observations and deduction.*

Observations

1. On adding barium chloride solution, a white precipitate is formed.
2. The white precipitate does not dissolve in hydrochloric acid.

Deduction

The solution of B10 in distilled water contains sulphate ions, SO_4^{2-}, which form a white precipitate of barium sulphate, $BaSO_4$, with barium chloride.

(ii) To the other part of the solution of the B10 in distilled water add lead nitrate solution. Record your observation and deduction.

Observation

A thick, white precipitate is formed on adding lead nitrate solution.

Inference

The solution contains SO_4^{2-} ions which form a white precipitate of lead sulphate, $PbSO_4$, when lead nitrate solution is added.

(c) Write down the name and the formula of the substance B10.

Name of B10

The substance B10 is ammonium sulphate.

Formula of B10

The formula of B 10 is $(NH_4)_2SO_4$.

Exercise 12

(a) Take a small portion of substance B25 in a test tube. Add 4 cm³ of distilled water and shake well. Allow the mixture to stand for about 5 minutes. Filter, and test the filtrate with litmus papers. Note down your observations and deduction.

Observations

The filterate has no effect on blue litmus paper but it turns red litmus paper blue.

Deduction

The solution of substance B25 in water is alkaline in nature.

(b) In a clean test tube take about 5 cm³ of dilute hydrochloric acid. Add substance B25 in small lots with vigorous shaking after each addition. Stop adding when some of it remains undissolved and settles down undissolved. Filter, and divide the solution in two parts.

(i) To one part of the filtrate add sodium hydroxide, a few drops at a time, till it is in excess.

(ii) To the other portion add sodium carbonate solution. Record your observations and inference about the metal ion present in the filtrate explaining the observations.

Observations

1. On adding sodium hydroxide solution a white precipitate is formed. The precipitate consists of white particles and does not dissolve in the excess of alkali.
2. On adding sodium carbonate solution a thick, white precipitate is formed.

Deduction

The metal ion present is Ca^{2+}. It forms a white precipitate of $Ca(OH)_2$ on adding sodium hydroxide and a thick, white precipitate of $CaCO_3$ on adding sodium carbonate.

(c) *Mix a pinch of B25 with about an equal quantity of ammonium chloride. Moisten the mixture with water, place in a test tube, and heat. Record your observations pertaining to the identification of the gas evolved. What inference can you draw about the identity of B25 from the tests performed?*

Observations

1. On heating a mixture of B25 and ammonium chloride, a colourless gas is evolved.
2. The gas has a characteristic pungent smell.
3. The gas turns moist red litmus paper blue.
4. When a glass rod moistened with concentrated hydrochloric acid is held near the mouth of the test tube, dense, white fumes are produced.

Identification of gas

The gas evolved is ammonia, NH_3.

Inference

The substance B25 forms an alkaline solution in water, displaces ammonia from ammonium chloride, and contains calcium ions, Ca^{2+}. Hence, it is calcium oxide, CaO.

Exercise 13

(a) *Take a small portion of the substance B28 in a clean test tube and dissolve it in the least quantity of water to make a concentrated solution. Add sodium hydroxide solution in small amounts till it is in excess. Write down your observations and deduction.*

Observations

1. On adding sodium hydroxide, a white precipitate consisting of fine particles is formed.
2. The precipitate does not dissolve in the excess of alkali.

Deduction

Substance B28 contains calcium, Ca^{2+}.

(b) *Take a pinch of B28 in a clean, dry test tube. Add 2cm³ of concentrated sulphuric acid and heat. Identify the gas evolved by suitable tests. Record your observations and deduction.*

Observations

1. When B28 is heated with concentrated sulphuric acid, a colourless gas is evolved.
2. The gas has a characteristic, pungent smell.
3. The gas turns moist blue litmus paper rod.
4. When a glass rod moistened with ammonia solution is held near the mouth of the test tube dense, white fumes are produced.

These observations show that the gas evolved is hydrogen chloride, HCl

Deduction

Substance B28 contains chloride ions, Cl^-, (or, substance B28 is a chloride).

(e) *Mix a pinch of B28 with about an equal quantity of manganese dioxide. Place the mixture in a dry test tube. Add about 2cm³ of concentrated sulphuric acid and heat. Identify the gas by suitable tests. Note down your observations and deduction.*

Observations

1. A greenish-yellow gas with an unpleasant, choking smell is evolved.
2. The gas turns moist blue litmus paper first red and then colourless.
3. The gas produces blue coloration in the starch-iodide paper held near the mouth of the test tube.

These observations show that the gas evolved is chlorine.

Deduction

Substance B28 contains chloride ions, Cl^-, (or, substance B28 is a chloride)

(d) *Name the substance B28 and write down its formula.*

Name of B28

Substance B28 is calcium chloride.

Formula of B28

Formula of substance B28 is $CaCl_2$.

Exercise 14

(a) *Take a portion of the substance B14 in a dry test tube and heat it carefully. Write down your observations and the inference based on them.*

Observations

1. On heating B14 colorless and odourless vapour is evolved.
2. The vapour condenses as drops of a colourless and odourless liquid on the upper part of the test tube.
3. The bright blue crystals of B14 change into a dirty white powder.

Inference

The substance B14 is a hydrated compound, probably hydrated copper sulphate.

(b) *Dissolve a small portion of B14 in water. To the solution add sodium hydroxide solution till it is in excess. Heat the mixture. Record your observations. Identify the metallic ion in B 14, giving reasons for your inference.*

Observations

1. A light blue precipitate is formed on adding sodium hydroxide solution.
2. On boiling, the light blue precipitate turns black.

Metallic ion identified

The metallic ion in B14 is Cu^{2+}. It forms blue precipitate of $Cu(OH)_2$ on treatment with alkali. When heated $Cu(OH)_2$ changes to black copper (II) oxide, CuO.

(c) Dissolve a small portion of B14 in water. To the solution add barium chloride solution and then dilute hydrochloric acid in excess. Write down your observations and the inference drawn.

Observations

1. On adding barium chloride solution to the solution of B14 in water, a white precipitate is formed.
2. The white precipitate does not dissolve in dilute hydrochloric acid.

Inference

The substance B14 contains sulphate radical, SO_4^{2-}, which forms a white precipitate of barium sulphate with barium chloride solution.

(d) Dissolve a pinch of B14 in water and put some clean iron filings in the solution. Allow to stand for 5-6 minutes, shaking occasionally. Record your observations.

Observations

1. Iron filings are coated with reddish brown particles of copper which get detatched on shaking.
2. The blue colour of the solution fades and finally changes to light green.

(e) Explain the changes observed in (d).

The experiments show that B14 is hydrated copper sulphate $CuSO_4.5H_2O$. Iron filings placed in its solution displace copper, and dissolve to form iron (II) sulphate. As a result the blue colour of the solution due to copper sulphate fades. When all the copper has been displaced, the solution contains iron (II) sulphate, $FeSO_4$ and is light green in colour.

APPENDIX II

Unsolved Exercises

Exercise 1

(*a*) Take about 1.0g of the substance A1 in a test tube. Add 2-3 cm^3 of concentrated hydrochloric acid, and heat gently. Filter, if necessary. Dilute the solution or the filtrate by adding about 10 cm^3 of water. Record your observations. What positive ion could be present in the solution ?

Keep the diluted solution for experiments (*b*), (*c*), and (*d*).

Observations

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Positive ion in solution

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(*b*) Take about 3cm^3 of the solution obtained in (*a*) in a test tube. Add sodium hydroxide solution till it is in excess. Heat the mixture. Write down your obsevations and inferences about the substances formed.

Observations

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Inference

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(*c*) Take about $3 cm^3$ of the solution obtained in (*a*) in a test tube. Slowly add ammonium hydroxide till it is in excess. Describe all that you would observe and make your deduction about the metal in A1.

Observations

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Deduction

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(*d*) Take about $3 cm^3$ of the solution obtained in (*a*) in a test tube and add a few pieces of iron filings. Allow to stand for 5-6 minutes. Record all your observations and draw inference based on them.

Observations

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Inference

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Note : *The substance A1 is copper oxide.*

Exercise 2

(*a*) Take a small amount of the substance A2 in a clean, dry test tube. Add a few drops of concentrated sulphuric acid, and heat. Write down your observations leading to the identification of the gas evolved. What inference can you draw from this experiment ?

Observations

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First test for the gas

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Second test for the gas

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Name of the gas

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Deduction

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(*b*) Take a little of the substance A2 in a clean test tube. Dissolve it in a small quantity of water. To the solution first add silver nitrate solution and then ammonium hydroxide solution. Describe your observations and draw inference about the substance A2.

Observations

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Deduction

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(*c*) Apply flame test to find out the metallic ion in A2. Describe how the test is performed. Note down the colour of the flame and the inference drawn.

Description of flame test

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Colour of the flame

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Inference

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(*d*) Name A2 and write its formula

Name and formula of A2

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Note : *A2 is potassium chloride.*

Exercise 3

Divide the substance A3 in four nearly equal parts.

(*a*) Take one part of A3 in a test tube and add about 2 cm^3 of sodium hydroxide solution. Heat the mixture and identify the gas evolved. Write all observations and your inference.

Observations

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Inference ....

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(*b*) Take another part of A3 in a dry test tube. Add a few drops of concentrated sulphuric acid and heat. Identify the gas evolved by suitable tests. Record your observations. What inference can you draw from this experiment ?

Observations

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Inference

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(*c*) Mix a part of A3 with about an equal quantity of manganese dioxide. Take the mixture in a test tube. Add a little concentrated sulphuric acid and heat. Apply suitable tests to identify the gas evolved. Write all observations.

Observations

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Gas identified

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(*d*) Dissolve the fourth part of A3 in a little amount of distilled water. To the solution obtained add silver nitrate solution. Divide the mixture in two parts. Add nitric acid to one part and ammonium hydroxide to the other one. Record all observations and the deduction made.

Observations

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Deduction

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(*e*) Name the substance A3 on the basis of earlier inferences.

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Note : *The substance A3 is ammonium chloride*

Exercise 4

(*a*) Take a pinch of the substance A4 in a dry, hard-glass test tube. Heat, first gently and the strongly. Record your observations. Identify the products other than the residue left behind, and make deductions.

Observations

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Products identified

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Deductions

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(*b*) Dissolve a small portion of A4 in distilled water in a clean test tube. Divide the solution in two equal parts.

1. To one part add sodium hydroxide solution, and then heat the mixture.

Observations

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Deduction

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II. To the second part add ammonium hydroxide slowly till it is in excess.

Observations

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Deduction

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(*c*) Take a pinch of A4 in a dry test tube. Add a little concentrated sulphuric acid and heat. Add a few copper turnings to the hot reaction mixture. Note down the observations. Identify the gas evolved and draw inference therefrom.

Observations

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Gas identified

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Deductions

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Note : *Substance A4 is hydrated copper nitrate.*

Exercise 5

(*a*) Take a pinch of the substance A5 in a clean test tube. Add 2 cm^3 of dilute hydrochloric acid, and gently heat. Identify the gas evolved. Note down your observations and inference.

Observations

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Gas identified ....

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Deduction ....

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(*b*) Take about 2 cm^3 of dilute sulphuric acid in a test tube and add to it a few drops of potassium permanganate solution. To the pink solution thus obtained add a pinch of A5, and gently heat. Write down your observations and the inference drawn therefrom.

Observations

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Inference ....

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(*c*) Dissolve a few crystals of ferric chloride in 1 cm^3 of water in a clean test tube. Add a few drops of concentrated sulphuric acid. Now add a pinch of A5., Heat to boiling. Cool the solution and add sodium hydroxide solution till in excess. Record your observations and explain the change

Observations

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Explanations of changes

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(*d*) Apply flame test to B14 to find out the metallic ion present. **Discribe the test.** Note down the colour of the flame and your deduction from it.

Description of flame test

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Colour of flame

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Name of the metallic ion in A5.

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Note : *The compound A5 is sodium sulphite.*

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